CHEMISTRY

For Class IX

Burette with Vinegar
Burette with Na OH
Clamp
Flask
iron stand

Mark showing known volume of solution
Meniscus
(a) (b) (c)

Sindh Textbook Board.

THE TEXTBOOK OF

CHEMISTRY

For
Class - IX

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PREFACE

The present day age is the age of science and technology. No nation can flourish without literacy particularly without scientific knowledge. Realising the developments in the science of chemistry which have taken tremendous strides, have involved, changes in the contents of the syllabus. Ministry of Education (Curriculum Wing), Islamabad, prepared a new syllabus of chemistry and got it approved by the government.

This book for class - IX has been prepared by the Sindh Textbook Board, Jamshoro in accordance with the new latest syllabus of chemistry. The Board has been able to secure the services of the competent teachers, who performed their duties with the best of their abilities and experiences in the field of chemistry.

In preparing this textbook, the authors have had in their mind a student of secondary level. At secondary level, the chemistry is begun before the student has acquired enough needed background in elementary mathematics for true understanding. However, mathematics is a part of the language of chemistry and a lack of familiarity with mathematics can become a barrier to success in understanding chemistry.

In order to increase the interest and the curiosity of the student at the secondary level; the subject matter has been made more effective by adding figures and diagrams and some points. Each chapter is divided into fairly short sections, using figures, tables, flow-diagrams to reinforce the text. S.I. units are used throughout the book.

Some of the salient features of this textbook are as follows:

(i) Basic concepts, relating to fundamentals of chemistry have been prepared in a logical, coherent and academic way. Historical backgrounds of chemistry are also introduced.

(ii) The topics are dealt, keeping in mind the secondary class students, to develop an interest in students in the study of chemistry as a discipline.

(iii) Every chapter begins with the introduction of the contents of the chapter, which are then dealt in the chapter with the wider context. Numerous worked out examples are introduced.

(iv) Each chapter concludes with a summary of the important facts and ideas to enable students to revise the contents of each chapter with ease.

(v) Review questions at the end in the exercise of each chapter provide interesting materials for tests, class discussions, revisions and homework.
Multiple-choice questions and fill in the blanks are also provided in the exercise of each chapter.
The latest curriculum, states the following learning objectives.
(a) To strengthen the concepts developed at secondary level and develop new concepts to provide a solid background for higher studies in chemistry.
(b) To acquaint and also to expose the students to different aspects of chemistry used in our daily life.

The book is written in simple language, and has clear concise and systematic presentation, with total commitment for a national cause.

Since there is always a room for improvement, so we apologize for any errors or omissions and welcome your constructive criticisms and positive suggestions which might improve the matter of book.

Editors

CHAPTER-1

بِسْمِ اللّٰهِ الرَّحْمٰنِ الرَّحِيْمِ

INTRODUCTION TO CHEMISTRY

You will learn in this chapter about :

* What is chemistry, its definition and importance.
* Land mark in the history of chemistry.
* Chemistry and society
* Branches of chemistry.
* The scientific approach in chemistry.

1.1 WHAT IS CHEMISTRY (Definition and its Importance)

People in the industrialized nations have a higher standard of living than the human race has ever known. More nutritious food, better health, better life and much more of this is due to chemistry. Chemistry enables us to design all sorts of materials, drugs to fight disease, pesticides to protect our health and fertilizers to grow our crops for abundant food, fuels for transportation, fibres to provide comfort and variety in clothes and a lot of other things.

Chemistry also helps us to understand the nature of our environment and about ourselves. The theories of chemistry illuminate our understanding of the material world from tiny atom to giant galaxies.

Definition:

Chemistry is the branch of science that deals with the properties, composition and structure of matter. Chemistry also deals with the changes involved in the matter. It also deals with the principles governing the changes which matter may undergo.

1.2 LANDMARKS IN THE HISTORY OF CHEMISTRY

Over the centuries, chemistry has undergone remarkable progress. Chemistry from the very beginning was, however, used in a number of

arts and crafts, such as pottery making, glass making, dyeing and in metallurgy. All these developments were empirical (experimental) that is they were achieved by trial and error and not on the basis of any systematic study.

1.2 (a) THE GREEK PERIOD

Greek philosophers like Plato (347 - 428 B.C), Aristotle (322-384 B.C), Democritus (357-460B.C) and many others contributed a lot in number of small ways to the early development of chemistry. They introduced the concept of elements, atoms and the chemical reactions. They thought that all matter was derived from four elements, earth, air, fire and water. These elements have properties of their own such as hot, dry, cold and wet. According to them, fire was hot and dry, earth was dry and cold, water was cold, hot and wet and air was cold, wet and hot.

The Romans developed the chemical arts still further. They improved metallurgical processes and introduced the enamelling of pottery. However, they developed little theoretical knowledge in this regard. Their works were all empirical.

(b) THE MUSLIM PERIOD

The Muslim period was from 600 to 1600 A.D in the history of chemistry and is known as the period of alchemists. In the middle ages, chemistry was given a spurt of work. During the first few centuries of Hijra, the Muslim scientists made rich contributions to the various branches of science, specially in the field of chemistry and introduced scientific methods and experimentations. The modern scientific knowledge is based on the contributions of these Muslim scholars.

The alchemists developed and used many laboratory equipments such as funnels, beakers, crucibles for melting and fusion, retorts for distillation, balances for weighing, etc. They discovered various acids, alcohols and medicines.

Jabir Ibne-Haiyan (721 - 803 A.D), generally known as the father of alchemy, invented experimental methods for the preparations of nitric acid, hydrochloric acid and white lead. He also developed methods for the extraction of metals from their ores and dyeing cloths.

Al-Razi (862 - 930 A.D) was a physician, alchemist and a philosopher. He was an expert surgeon and was the first to use opium as an anaesthesia. He divided the substances into living and non-living origins, which was later adopted by Berzellius, in 1806 to classify chemical compounds on the basis of their origins as organic and inorganic compounds. Al - Razi prepared

ethyl alcohol by the fermentation process.

Al - Beruni (973 - 1048 A.D) contributed a lot in physics, metaphysics, mathematics, geography and history. In the field of chemistry, he determined the densities of different substances.

Ibne-Sina (980 - 1037 A.D) was famous for his contribution in the field of medicines, medicinal chemistry, philosophy, mathematics and astronomy.

These Muslim alchemists were interested more in finding a way to prolong life and to convert baser metals like lead, copper into gold. Although their efforts were futile but their researches led to the discoveries of many substances and laid the foundation of chemistry. By the late sixteenth century, a lot of chemical laboratories sprang up and the important reagents like sulphuric acid, nitric acid, hydrochloric acid, silver nitrate, etc were discovered. Chemists of that period, however, devoted their energies mainly to the production of drugs for the use of medicines.

(c) THE MODERN PERIOD

Robert Boyle (1627 - 1691 A.D) described affectionately as the father of modern chemistry, was the first to put forward the idea that chemistry should be regarded as a systematic investigation of nature with the sole aim of promoting knowledge. As a result, lot of discoveries were made during later years.

J. Black (1728 - 1799) made a study of carbon dioxide.

J. Priestly (1733 - 1804) discovered oxygen, sulphur dioxide and hydrogen chloride.

Scheele (1742 - 1786) discovered chlorine. Cavendish (1731-1810) discovered hydrogen, Lavoisior (1743 - 1794) discovered that oxygen constituted about one-fifth of air. John Dalton (1766 - 1844) made a great stride in chemistry with his atomic theory of matter. The idea of atoms and molecules became the established fact and the concept of atomic weight was also given by John Dalton.

Gay - Lussac (1778 - 1850), Avogadro (1776 - 1856), Dulang (1785 - 1838) and Petit (1741 - 1820) and others led to the determination of relative atomic and molecular masses of many substances. Meanwhile J. J. Berzelius (1779 - 1848) introduced the idea of symbols, formulae and chemical equation to make the study more systematic. This was further enhanced by Mendeleeff (1824 - 1907) who discovered the periodic arrangement of the elements. The ionic theory and the laws of electrolysis were put forward by Arrhenius (1859 - 1927) and M. Faraday (1791 - 1867), respectively. The works of J.J. Thomson (1856 - 1940), H. Becquerel (1852 - 1908), M. Curie (1867 - 1934), Neil Bohr (1885 - 1962) and E. Rutherford (1891 - 1937) led

to the discoveries of structure of atom, radioactivity, there by opening the door of the present nuclear age.

1.3 CHEMISTRY AND SOCIETY

There are three significant reasons to study chemistry. First, chemistry has important practical applications in the society. The development of life saving drugs is one and a complete list would touch upon most areas of modern technology.

Second, chemistry is an intellectual enterprise, a way of explaining our material world.

Finally, chemistry figures prominently in other fields, such as in biology, in the advancement of medicines. Chemistry in every field is a useful intellectual tool for making important decisions.

The role of chemistry in the prevailing society is of enormous benefits. We are familiar with many chemicals which have become part and parcel of our daily life.

Chlorine has become an important commercial chemical. Today chlorine has turned into the main element for producing more than one thousand chlorine compounds, which are of great importance in chemical industry, producing Poly Vinyl Chloride (PVC) as plastics for pipes. Other Chlorine compounds are used as bleaching agent, disinfectants, solvents, pesticides, refrigerants, flame retardant and drugs.

Cholera, typhoid fever and dysentery are dangerous diseases, transmitted through impure drinking water. All such water - borne diseases are all but eliminated from most of the parts of the world, when water supplies are treated with chlorine which kill pathogenic (disease-causing) organisms. Chemistry plays an important role in the modern world. For example food, synthetic fibers, plastics, medicines, soap, detergents, cosmetics, cement, fertilizers, glass and explosives are the major gifts of chemistry.

The applications of chemistry are much like the science itself, undergoing constant changes. Chemistry is an extremely practical science and has deep influence on our daily living. It matters with the protection of environment, providing our everyday needs of food, clothing and shelters. Using chemistry we have discovered pharmaceutical chemicals that enhance our health and prolong our lives. Unfortunately some chemicals have the potentials of harming our health or the environment. It is in our best interest as educated citizens and consumers to understand the profound effect, both positive and negative that chemicals have on our lives.

We use fluoride compounds such as SnF_2, Na_2PO_4. F (Sodium fluoro phosphate) and NaF in our tooth pastes to protect and control tooth decay. It is a great beneficence of chemistry on the society.

1.4 BRANCHES OF CHEMISTRY

There are various branches of chemistry. The main branches of chemistry are:

(1) Physical chemistry
(2) Organic chemistry
(3) Inorganic chemistry
(4) Analytical chemistry
(5) Bio chemistry
(6) Industrial and Applied chemistry
(7) Nuclear chemistry
(8) Environmental chemistry
(9) Polymeric chemistry

(1) Physical Chemistry:

It is the branch of chemistry that deals with the laws and the principles governing the combination of atoms and molecules in chemical reactions.

(2) Organic Chemistry:

It is the branch of chemistry that deals with carbon compounds with the exception of CO_2, CO, metal carbonates, bicarbonates and carbides. Actually it is the chemistry of hydrocarbons and their derivatives. Most of the consumers products are organic in nature.

(3) Inorganic Chemistry:

It is the branch of chemistry that deals with the chemistry of elements and their compounds, generally obtained from non-living organisms i.e. from minerals.

(4) Analytical Chemistry:

It is the branch of chemistry that deals with the study of the methods and techniques involved to determine the kind, quality and quantity of various components in a given substance.

(5) Biochemistry:

It is the branch of chemistry that deals with the compounds of living organisms i.e. plants and animals and their metabolism in the living body.

(6) Industrial Chemistry:

It is the branch of chemistry that deals with the study of different

chemical processes involved in the chemical industries for the manufacture of synthetic products like glass, cement, paper, soda ash, fertilizers, medicines etc.

(7) Nuclear Chemistry:

It is the branch of chemistry that deals with the study of changes occurring in the nuclei of atoms, accompanied by the emission of invisible radiations.

(8) Environmental Chemistry:

It is the branch of chemistry that deals with the study of the interaction of chemical materials and their effect on the environment of animals and plants. Personal hygiene, pollution and health hazards are important areas of environmental chemistry.

(9) Polymeric Chemistry:

It is the branch of chemistry that deals specially with the study of polymerization and the products obtained through the process of polymerization such as plastics, synthetic fibers, papers etc.

Every branch of chemistry has its own importance in human life. Biochemistry is the backbone of medical sciences. Environmental chemistry tells us about the environment composition and that how one can protect its environment from environmental hazards. Industrial chemistry helps us in the manufacturing of the industrial products and their uses. Analytical chemistry is important to understand the composition of compounds, quality of the products, analysis of biological samples, such as water, milk, urine, blood, soil and the use of research techniques, such as chromotography and spectroscopy. Nuclear chemistry provides us radio isotopes for the treatment of many diseases, such as cancer and also to give atomic energy for the benefits of mankind.

1.5 THE SCIENTIFIC APPROACH IN CHEMISTRY (Observation, Hypothesis, Theory and Law)

Over the years science has developed through a series of discoveries which started off as observed natural phenomenon which had to be explained. This was done by using scientific method in a systematic manner.

(1) Observation:

Different people observe a phenomenon in different ways. Some of us observe something very critically to extract from it a new point.

Observation is a basic tool to go forth for elaborating a phenomenon but it may vary from person to person according to his own skill of elaboration. Observation of a thing is one of the scientific approaches in chemistry.

(2) Hypothesis:

When a phenomenon is observed; a scientist ponders over it and carries out relevant experiments. He sieves through the data and arrives at a possible explanation for the nature of the phenomenon. This explanation; which is still only a trial idea is called hypothesis. It may or may not undergo a change as a result further investigations and accumulation of more knowledge or facts.

(3) Theory:

The scientist conveys his hypothesis to other workers of the same field for the discussion and for further experimentation. When the hypothesis is supported by a large amount of different types of observations and experiments, then it becomes a theory i.e. scientifically acceptable idea or principle to explain a phenomenon. A good theory predicts new facts and unravels new relationship between naturally occurring phenomenon.

(4) Scientific Law:

A theory which is tested again and again and found to fit the facts and from which valid predictions may be made is then known as scientific law or principle.

However, not all hypothesis and theories pass successfully to become scientific laws. Some may sound very convincing and be supported by mathematical calculations but are very difficult to prove experimentally. This is invariably due to the material under investigation or the lack of suitable working equipments. A typical example is Avogadro's hypothesis which has not been proved conclusively and yet it is accepted as Avogadro's law.

Science cherishes all forms of ideas and proposals. Even obsolete ideas are kept as reference. It is said that there is no end to knowledge, so development in science too may have no limits.

SUMMARY

1. Chemistry is the branch of science that deals with the properties, composition and structure of matter. It also deals with the principles and laws governing the changes involved in the matter.
2. Chemistry is as old as human civilization. Over the centuries chemistry has undergone remarkable progress and was used in pottery making, glass making, dyeing and in metallurgies.
3. Greek philosphers (347 to 460 B.C) Plato, Aristotle, Democritus and many others contributed in the early developments of chemistry in a number of ways. The concepts about elments, atoms and chemical reactions were introduced by them. However they developed little theoritical knowledge in chemistry. Their works were mostly empirical.
4. The Muslim period started from 600 A.D and ended about 1600 A.D in the history of chemistry and was known as the period of alchemia. During the first four centuries of Hijra, the Muslim scientists contributed richly towards many branches of science, specially in the field of chemistry and introduced scientific methods and experimentations. The alchemists developed and used many laboratory equipments, such as funnels, beakers, crucibles, retorts for distillation, balances etc. The modern scientific knowledge is based on the contributions of these Muslim scientists or scholors like Jabir-Ibne-Haiyan, Al-Razi, Al Beruni, Ibne-Sina, etc.
5. Modern chemistry began in the 17th and 18th century as a result of experimentations, free discussion and communication of work of chemists through out the world. Robert Boyle is described as the father of modern chemistry. J.J. Berzelius introduced the idea of symbols, formulae and chemical equations to make the study of chemistry more systematic. J.J Thomson, Henry Becqueral, Rutherford, Neil Bohr, M.Curie led to the discoveries of structure of atom, radioactivity which opened the door of present nuclear age.
6. Chemistry has great impact on the society. It has important practical applications in the society, in medicines in modern technology, industries, environmental composition and its protection from environmental hazards.
7. There are various branches of chemistry which are physical chemsitry, organic chemistry, inorganic chemistry, analytical chemistry, biochemsitry, industrial chemistry, environmental chemistry, nuclear chemistry, etc.

8. The observed natural phenomenon had to be explained by different scientific methods in systematic manner. The scientific approach in chemistry can be described through observations, hypothesis, theory and law.

Observation is the basic tool for elaborating a phenomenon but it may vary from person to person according to his own skill. It is one of the scientific approaches in chemistry.

After observation of a phenomenon, scientist ponders over it and carries out relevant experiments and arrives at a possible explanation for the nature of the phenomenon. This explanation which is a trial idea is known as hypothesis. A hypothesis is liable to undergo changes as a result of further investigations.

When the hypothesis is supported by a large number of different types of observations and experiments by other workers on the same phenomenon, then a theory is developed. A good theory predicts new facts and figures, unreavelling new relationships between naturally occurring phenomenon.

When a theory is tested again and again and is found fit according to the facts, giving valued predictions, is called as scientific law.

EXERCISE

1. Fill in the blanks :

(i) The early Greeks believed that every thing in the universe was made up of four elements _____ ; _____ , _____ and _____.

(ii) Al-Razi divided chemical substances on the basis of their ______.

(iii) _________ is the branch of chemistry which deals with the carbon compounds.

(iv) Biochemistry is the backbone of _________.

(v) PVC which is a plastic is the short name for ________.

(vi) Oxygen was discovered by _________.

(vii) The best disinfectant is _________.

(viii) The periodic arrangement was the result of __________ work.

2. **Write answer of the following questions :**

(i). Write a note on the historical development of chemistry with special mentioning of the contribution of Muslim scientists in the field of chemistry.

(ii). Define chemistry and describe its importance.

(iii). What important role chemistry plays in the society ?

(iv) Name the different branches of chemistry and define them.

(v). What do you mean by the scientific approach in chemistry? How will you differentiate between hypothesis and theory ?

(vi). What is scientific law?

CHAPTER-2

CHEMICAL COMBINATIONS

You will learn in this chapter about :

* Laws of chemical combinations.
* Atomic mass.
* Chemical formula.
* Empirical formula.
* Molecular formula.
* Molecular mass and formula mass.
* Mole.
* Avogadro's number.
* Chemical reactions and chemical changes.
* Chemical reactions and types of chemical reactions.
* Chemical equations.
* Writing of chemical equations.
* The balancing of chemical equations.
* Concept of mole ratios and balanced chemical equations.
* Calculations based on chemical equations.

2.1 LAWS OF CHEMICAL COMBINATIONS

Introduction:-

Chemistry deals with the matter and the changes occurring in it, Chemists are particularly interested in those changes, where one or more substances are changed into quite different substances. They had found that these chemical changes are governed by some empirical laws known as laws of chemical combinations.

These laws are:

1. Law of conservation of mass

2. Law of constant composition (or)
   Law of definite proportions
3. Law of multiple proportions
4. Law of reciprocal proportions.

1. Law of Conservation of Mass:

Matter under goes changes. However, it has been found that in all chemical changes, there is no change in the mass of the substances being changed. For example, in iron (Fe) increase in weight on rusting is because of its combination with oxygen from the air and the increase in weight is exactly equal to the weight of oxygen combined. The French Chemist Lavosier, (1785) tried to learn about chemical changes by weighing the quantities of substances used in chemical reactions. He found that when a chemical reaction was carried out in a closed system, the total weight of the system was not changed. The most important chemical reaction that Lavosier performed was the decomposition of the red oxide of mercury to form metallic mercury and a gas, he named this gas as oxygen. Lavosier summarised his findings by formulating a law, which is known as **law of conservation of mass**. It states that mass is neither created nor destroyed during a chemical reaction. In other words , *In any chemical reaction the initial weight of reacting substances is equal to the final weight of the products.*

The law of conservation of mass may be demonstrated by the union of hydrogen (H_2) and oxygen (O_2) to form water. If the H_2 and O_2 are weighed before they unite, it will be found that their combined weight is equal to the weight of water (H_2O) formed. This is shown below in fig.2.1.

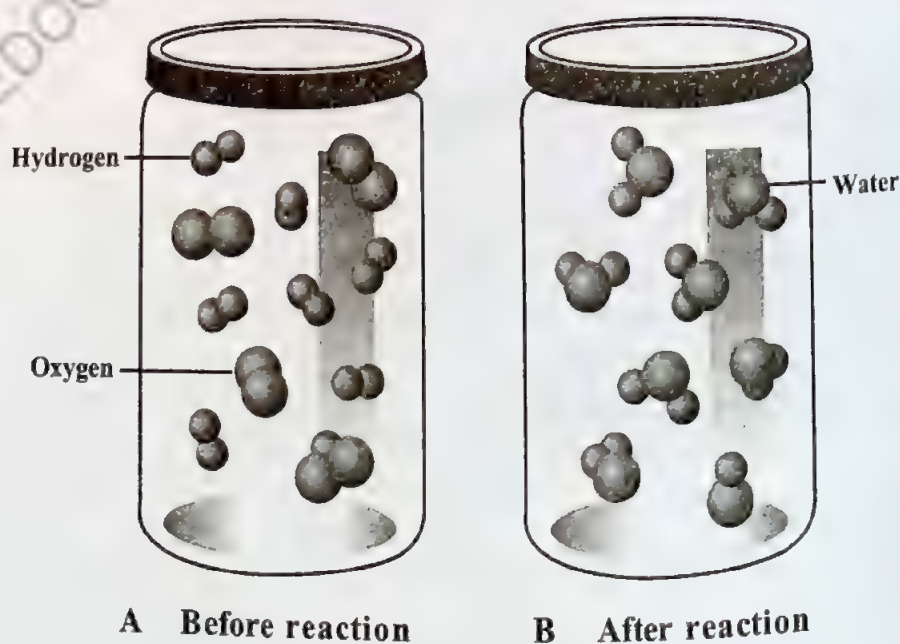
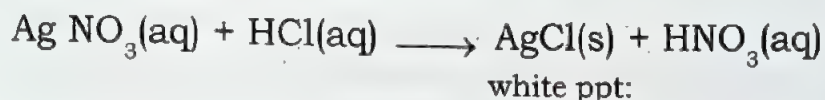



Fig: 2.1

Practical Verification: (Landolt Experiment)

German Chemist H. Landolt, studied about fifteen different chemical reactions with a great skill, to test the validity of the law of conservation of mass. For this, he took H-shaped tube as shown in fig. 2.2 and filled the two limbs A and B, with silver nitrate (Ag NO_3) in limb A and hydrochloric acid (HCl) in limb B. The tube was sealed so that the material could not escape outside. The tube was weighed initially in a vertical position so that the solutions should not intermix with each other. The reactants were mixed by inverting and shaking the tube. The tube was weighed after mixing (on the formation of white precipitate of (AgCl). He observed that weight remains same.

$$Ag\,NO_3(aq) + HCl(aq) \longrightarrow \underset{\text{white ppt:}}{AgCl(s)} + HNO_3(aq)$$

Thus total mass of the substance before the reaction is equal to the total mass of the substances after the reaction.

A B

$AgNO_3$ HCl

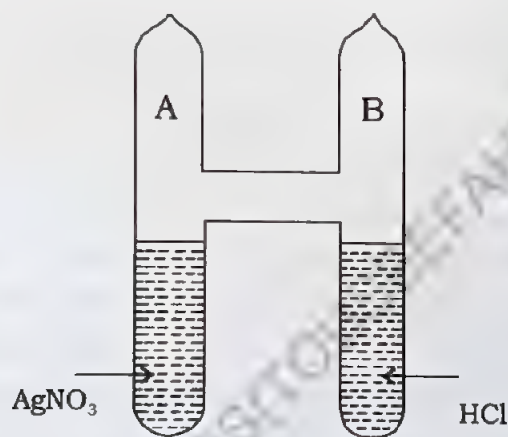



Fig: 2.2

In ordinary chemical changes, relatively small amount of energy is released. But in nuclear changes where uranium atoms undergo fission (break up) into smaller atoms plus neutrons, the total mass of products is noticeably less than that of starting material. This clearly indicates that some mass of uranium has been converted into energy, which is evident to us as heat and radiation.

The relationship between mass that is lost and the energy that is released is given by the equation.

$$E = mc^2$$

Where (E) is the energy in ergs, (m) is the mass in grams and (C) is the velocity of light in centimeters per second, ($3x10^{10}$ cm/sec). This relationship between mass and energy was first proposed in (1906), by the famous Physicist and Mathematician, Albert Einstien.

It follows that for every chemical change, there will be a mass change.

But this mass change is too small that no one has yet been able to detect through weighing techniques.

Hence the law of conservation of mass is, therefore, still valid from practical view point for ordinary chemical reactions i.e., *"there is no detectable gain or loss of mass in a chemical reaction"*.

2. Law of Constant Composition or Law of Definite (fixed) Proportion

By the end of Eighteenth century, chemists showed that a given compound has a definite (constant) composition. French chemist Louis Proust in (1799) summarized this result in the form of the law of definite proportion (also known as constant composition) which states, that different samples of the same compound always contain the same elements combined together in the same proportions by mass.

For instance every sample of pure water, though prepared in the laboratory or obtained from rain, river or water pump contains one part hydrogen (H) and (8) parts oxygen (O) by mass

eg: H_2O
2:16
1:8 (parts by mass)

One of the earliest illustrations of the law of definite proportions is found in the work of Swedish chemist J. J Berzelius (1779-1848).

Berzelius heated 10g of lead (Pb) with various amounts of sulphur (S). He got exactly 11.56g of lead sulphide and the excess of sulphur was left over, when he used 18g of lead (Pb) with 1.56g of sulphur (S), he got exactly 11.56g of lead sulphide (PbS) and the 8g of lead (Pb) remained unused.

These reactions are shown diagrammatically in Fig. 2.3

10g of lead + 1.56g of sulphur → 11.56g of lead sulphide

18g of lead + 1.56g of sulphur → 11.56g of lead sulphide + 8g of lead

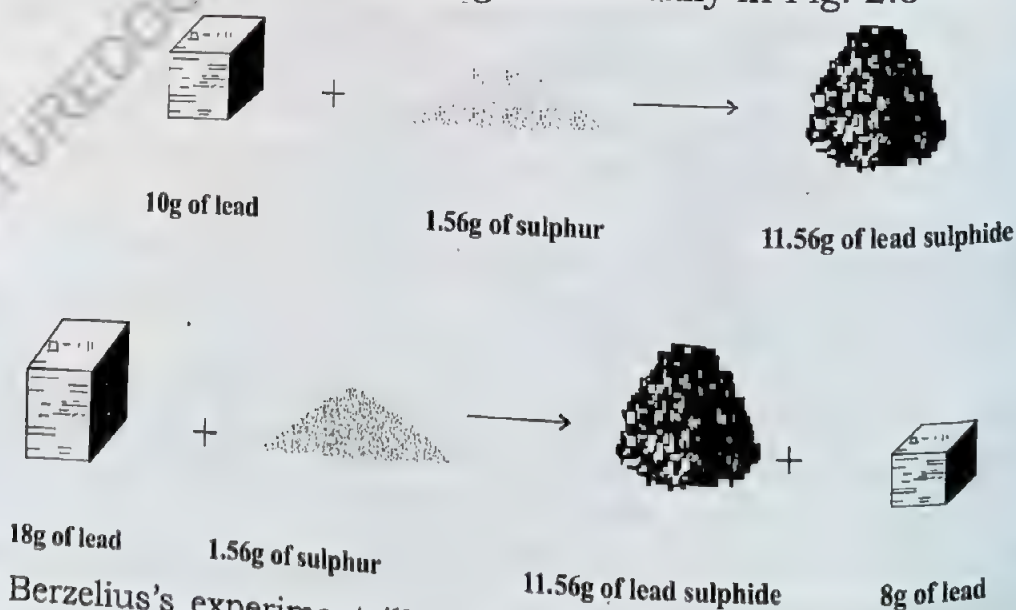



Fig. 2.3 Berzelius's experiment illustrating the law of definite proportions.

3. Law of Multiple Proportion:

The fact that the same element, can combine in more than one ratio to form different compounds was published by, John Dalton, (1803) in the form of law of multiple proportion. "*It states that if two elements combine to form more than one compounds, the masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers or simple multiple ratio.*"

For example: Carbon (C) forms two stable compounds with oxygen (O) namely carbon monoxide (CO) and carbon dioxide (CO_2).

Compound	Mass of Carbon (C)	Mass of Oxygen (O)	Ratio of Oxygen (O)
Carbon monoxide CO	12	16	1
Carbon dioxide CO_2	12	32	2

The different masses of oxygen 16 and 32 which combine with the fixed mass of C (12g) are in ratio of [16:32], i.e. 1:2, which is simple whole number ratio, and obeys the law of multiple proportion.

Another illustration of this law is the formation of water (H_2O) and (H_2O_2) from hydrogen and oxygen.

Compound	Mass of Hydrogen (H)	Mass of Oxygen (O)	Ratio of Oxygen (O)
Water H_2O	2	16	1
Hydrogen peroxide (H_2O_2)	2	32	2

The different masses of oxygen 16:32, which combine with the fixed mass of hydrogen (2g) are in ratio of 16:32 i. e. 1:2 which is again in a ratio of simple whole numbers.

The excellent illustration of law of multiple proportion is furnished, when the elements nitrogen (N) and oxygen (O) combine together to form a series of five oxides of nitrogen, in which these two elements are present in different proportions.

S.No.	Name of Oxides	Mass of (N)	Mass of (O)	Fixed mass of (N)	Variable mass of (O)	Ratio of (O)
1.	Nitrous oxide (N_2O)	28	16	14	8 (1x8)	1
2.	Nitric oxide (NO)	14	16	14	16 (2x8)	2
3.	Nitrogen trioxide (N_2O_3)	28	48	14	24 (3x8)	3
4.	Nitrogen tetra oxide (N_2O_4)	28	64	14	32 (4x8)	4
5.	Nitrogen penta oxide (N_2O_5)	28	80	14	40 (5x8)	5

By fixing the mass of (N), the mass of (O) in different oxides varies.

i.e. 8 : 16 : 24 : 32 : 40
1 : 2 : 3 : 4 : 5

These figures (in multiple ratios), are according to the law of multiple proportion.

4. Law of Reciprocal Proportion:

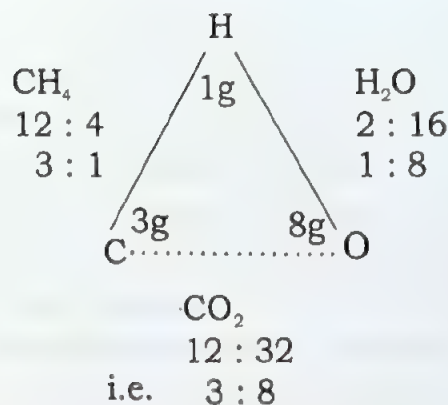
This law was enunciated by Ritcher in (1792-94). *It states that "when two different elements separately combine with the fixed mass of third element, the proportions in which they combine with one another shall be either in the same ratio or some simple multiple of it".*

For instance, when two elements C and O separately combine with H to form methane (CH_4) and water (H_2O) respectively it is very clear, that in methane 3g of C combine with 1g of hydrogen, and in water (H_2O) 8g of O combine with the same (fixed) mass i.e (1g) of H. Now when C and O combine with each other to form carbon dioxide (CO_2), they do so in the same proportion i. e. 12:32 = 3:8 parts by mass.

H

CH_4 12 : 4 3 : 1 — 1g — H_2O 2 : 16 1 : 8

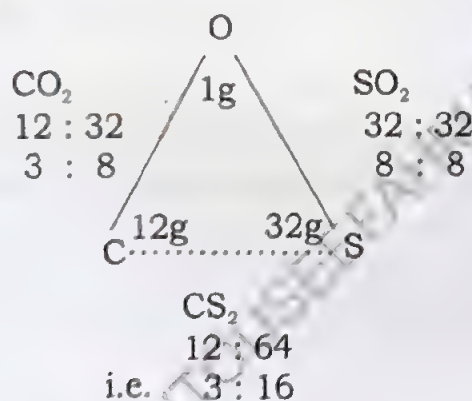
C — 3g .... 8g — O

CO_2
12 : 32
i.e. 3 : 8

Another illustration of law of reciprocal proportion is provided when, 12g of C combine with 32g of O to form carbon dioxide (CO_2) and 32g of sulphur (S) combine with the same (fixed) mass of oxygen (O) i. e. 32 g to form sulphur dioxide.

O

CO_2 12 : 32 3 : 8 — 1g — SO_2 32 : 32 8 : 8

C — 12g .... 32g — S

CS_2
12 : 64
i.e. 3 : 16

The above example, shows that the mass of C and S that combine with the same mass of O are in the proportion of 12:64 i. e. 3:16.

According to the statement of law of reciprocal proportion, the proportion in which C and S combine with one another shall be either in the same ratio (3:8) or some simple multiple of it i. e (3:16).

It is very clear that in the formation of carbon disulphide (CS_2), C and S combine in the ratio of (12:64) i.e (3:16) which is simple multiple of (3:8).

2.2 ATOMIC MASS

The mass of an atom depends upon the number of protons and neutrons present in it. As the atoms are extremely small particles, it is difficult to weigh them directly. For example the mass of single hydrogen (H) atom, is 1.6×10^{-24}g (0.000 000 000 000 000 000 000 0016g). Clearly we cannot weigh a hydrogen atom or any other kind of atom, by placing it

on a balance pan. Scientist need special method to obtain the mass of an atom by comparison to a standard mass.

In 1961, by an international agreement, an atom of C-12, that has 6 protons and 6 neutrons has a mass of exactly 12 atomic mass units (a.m.u.) is taken as a standard. So one atomic mass unit (1 a.m.u.) is defined as a mass exactly equal to one-twelfth the mass of C-12 atom. Since most elements consist of a mixture of isotopes (Isotopes are the atoms of same element, having same atomic number but different atomic masses). For example, naturally occurring carbon is composed of 98.889 percent C-12 and 1.111 percent C-13. Thus average atomic mass of C-atom becomes 12.011 a.m.u.

The atomic mass of an element is now taken as, the average mass of natural mixture of isotopes which is compared to the mass of one atom of C-12 a.m.u.

Thus the atomic mass of oxygen (O)=16 a.m.u. and that of sulphur (S) = 32 a.m.u.

2.2.1 Empirical Formula and Molecular Formula: (E.F and M.F)

A formula is a combination of symbols for atoms or ions, that are held together chemically in a compound. By formula we mean not only the elements present but also ratios in which the atoms are combined. Hence we will discuss two types of formulas i.e. Empirical formula and Molecular formula.

2.2.1(a) Empirical Formula (E.F): (Simplest formula).

A formula that gives only the relative number of each type of atoms present in a molecule. In other words, the empirical formula does not necessarily give the actual number of atoms in a molecule. For example, the molecular formula of benzene is C_6H_6. This formula indicates that benzene molecule consists of (6) carbon atoms and (6) hydrogen atoms. The ratio of carbon (C) to hydrogen (H) atoms in this molecule is 6:6 or 1:1. The empirical formula of benzene is, therefore written as (CH).

Thus empirical formula tells us which elements are present and their simplest atomic ratio, but not necessarily the actual number of atoms present in the molecule.

Consider another example, the molecule of glucose ($C_6H_{12}O_6$) in which the ratio of C, H and O atoms is 6:12:6 i.e. 1:2:1. The empirical formula of glucose is, therefore, (CH_2O).

2.2.1 (b) Molecular Formula (M. F):

Molecular formula indicates the actual number and type of atoms in a molecule. It can either be same as empirical formula or some simple multiple of it. Mathematically, M. F = (E.F) n, where (n) is the whole number. For example the molecule of CO_2 consists of one atom of carbon in combination with two atoms of oxygen. The formula (CO_2) is the molecular formula of carbon dioxide. It represents the true composition of a molecule of the compound. The molecular formula may be same as empirical formula as in the case of CO_2 or some simple multiple of empirical formula. Thus the molecule of glucose which shows that the molecule of glucose, consists of (6) carbon, (12) hydrogen and (6) oxygen atoms and its simplest atomic ratio i.e empirical formula is (CH_2O). Thus the molecular formula of glucose ($C_6H_{12}O_6$) is equal to $(CH_2O)_6$ or six times to empirical formula (CH_2O).

It follows that molecular formula of glucose is six times the empirical formula, which is obtained by M. F = $(E.\ F)_n$ where (n) is the whole number, and in glucose n=6.

Mathematically,

$$n = \frac{\text{M.F. weight}}{\text{E.F. weight}}$$

For many molecules, the molecular formula and empirical formula are the same, some examples are formaldehyde (CH_2O), ammonia (NH_3) and methane (CH_4).

2.2.2 Molecular Formula Mass:

The molecular formula mass (molecular mass) of a substance is the sum of the atomic masses of all atoms present in the molecular formula of a substance or molecule. Taking as an examples, let us calculate the molecular formula mass of CO_2. The molecule of CO_2 contains one atom of C and two atoms of O. The atomic masses of C and O are 12 a. m. u. and 16 a. m. u, respectively.

C	= 12 x 1	=	12 a.m.u.
O	= 16 x 2	=	32 a.m.u.
Molecular formula mass of CO_2		=	44 a.m.u.

For example, Compute the molecular formula mass of ozone (O_3). Molecular formula mass is calculated by adding together the atomic masses of the

constituent atoms. The ozone (O_3) molecule contains three oxygen atoms each of which has mass equal to 16a.m.u.
Therefore molecular formula mass of ozone (O_3) = 3 x 16 = 48 a.m.u.

2.2.3 Formula Mass:

Formula mass of substances is the sum of the atomic masses of all atoms in a formula unit of the substance. For example, we can calculate the formula mass of sodium chloride (NaCl), a common salt, by adding the atomic masses of all atoms in the formula unit, expressed in (a.m.u.). the atomic masses of (Na) and (Cl) are 23 a.m.u. and 35.5 a.m.u., respectively.

Na = 23 x 1 = 23 a.m.u.
Cl = 35.5 x 1 = 35.5 a.m.u.
∴ Formula mass of NaCl = 58.5 a.m.u.

Remember

That the term molecular mass applies to molecular compounds. The term formula mass can be used with either molecular compounds or ionic compounds. The term molecular mass can not be used with ionic compounds because there are no discrete molecules in ionic compounds.

2.2.4 Molar Mass:

Molar mass of a substance is its relative molecular mass expressed in grams. Thus molar mass of a substance has a fixed unit. For example, 1 mole of (C) is equal to its atomic mass expressed in grams.

Molecular mass of C = 12 a.m.u. and therefore the molar mass of carbon would be 12g.

For example, calculate the molar mass of ammonia (NH_3).
The molar mass, is obtained by adding the atomic masses of component atoms.

N = 1 x 14g = 14g
H = 3 x 1g = 3g
Molar mass of NH_3 = 17g

Remember, that relative molecular mass of NH_3 = 17 a.m.u.

2.3 MOLE

(A practical chemical unit for handling atoms and molecules).
Since atoms and molecules are so small, it is impossible to handle and

count atoms and molecules individually. Therefore, the Chemists devised a special unit to describe very large number of atoms, ions and molecules.

It is called mole and is abbreviated as mol. A mole can be defined as *"the molecular mass, atomic mass and formula mass of a substance expressed in grams. Thus, 12g of carbon is equal to 1 mol of carbon atoms. 24g of C is equal to 2 mol of carbon atoms"*. The mole concept tries to give a practical meaning to the mass of reactants and products in chemical reactions in terms of the number of particles (atoms, molecules or ions) involved. The actual countable number of particles in one mole of a substance is 6.02×10^{23} particles, which is referred as Avogadro number.

The (S.I.) definition of mole is the amount of substance, containing as many elementary particles (units) as there are atoms in exactly 12g of C-12 a.m.u.

It is also defined as the mass of any substance equal to its atomic mass, molecular mass or formula mass in grams.

Thus.

1 mole of C	=	12 g
1 mole of Mg	=	24 g
1 mole of H_2O	=	18 g
1 mole of CO_2	=	44 g
1 mole of $CaCO_3$	=	100g
1 mole of Fe_2O_3	=	160g

$$\text{By formula, number of moles} = \frac{\text{Given mass of substances}}{\text{Atomic mass or formula mass}}$$

As, formula mass, represents the both covalent and electrovalent compounds.

2.3.1 Avogadro's Number (N_A):

A mole of substance always contains the same number of particles (*atoms, ions, molecules or formula units*) irrespective of its state, solid, liquid or gaseous, that is 6.02×10^{23} particles. This constant number has been determined by several methods, called Avogadro's number (symbol N_A),in the honour of Avogadro, the scientist who gave chemistry a method for finding atomic and molecular masses.

Thus.

1 mole of C	=	12 g	=	6.02×10^{23} atoms of carbon
1 mole of Mg	=	24 g	=	6.02×10^{23} atoms of magnesium
1 mole of H_2O	=	18 g	=	6.02×10^{23} molecules of water
1 mole of CO_2	=	44 g	=	6.02×10^{23} molecules of CO_2

1 mole of NaCl	=	58.5g	=	6.02×10^{23} F-units of NaCl
1 mole of $CaCO_3$	=	100g	=	6.02×10^{23} F-units of $CaCO_3$
1 mole of Na^+	=	23g	=	6.02×10^{23} ions of Na^+
1 mole of Cl^-	=	35.5g	=	6.02×10^{23} ions of Cl^-

Conversion of Mass into Moles and Moles into Mass of Substance

Problem 1. Calculate the number of moles, in 50 g of each.

(a) Na (b) H_2O

Solution:

$$\text{Number of moles} = \frac{\text{Given mass of substance}}{\text{Atomic mass or Formula mass}}$$

(a) Given,

i) Number of moles = ?
ii) Given mass of Na = 50g
iii) Atomic mass of Na = 23 a.m.u.

$$\therefore \text{Number of moles of Na} = \frac{50}{23} = 2.173 \text{ moles of Na}$$

(b) Given,

i) Number of moles of H_2O = ?
ii) Formula mass of H_2O = 18 a.m.u.
iii) Given mass of H_2O = 50g

$$\therefore \text{Number of moles of } H_2O = \frac{50}{18} = 2.777 \text{ moles of } H_2O$$

Problem 2. What is the mass of 3 moles of each:

a) Al b) CO_2.

$$\text{Number of moles} = \frac{\text{Given mass of substance}}{\text{Atomic mass or Formula mass}}$$

$\therefore$ Mass of substance = number of moles x atomic mass or formula mass in grams

(a) Given,

i) Number of moles of Al = 3 moles
ii) Atomic mass of Al = 27 grams
iii) Mass of Al = Mole x atomic mass of Al
= 3 x 27=81g

(b) Given,

i) Number of moles of CO_2 = 3 moles
ii) Formula mass of CO_2 = 44g
iii) Mass of CO_2 = Mole x Formula mass of CO_2
= 3 x 44 = 132 g

Use of Avogadro's number:

a) To calculate the number of atoms or molecules in a given sample of substance.
b) To calculate the mass of single atom or molecule of any substance.

Problem: 1. Calculate the number of atoms in 9g of Al.

Solution: According to Avogadro's number.

1 mole of Al= 27g = 6.02×10^{23} atoms.

This shows that:

27 g of Al contain 6.02×10^{23} atoms of Al

1 g of Al will contain $\frac{6.02 \times 10^{23}}{27g}$

9g of Al will contain $\frac{6.02 \times 10^{23} \times 9}{27g}$ = 2.006×10^{23} atoms of Al.

This numerical can also be solved, by using the formula.

$$\text{Number of atoms} = \frac{N_A \times \text{Mass of substance}}{\text{Atomic mass}}$$

$$\text{Number of (Al)atoms} = \frac{6.02 \times 10^{23} \times 9g}{27g}$$

$$= 2.006 \times 10^{23} \text{ atoms of Al.}$$

Problem: 2. Calculate the number of molecules in 9g of CO_2.

Solution: According to Avogadro's number

1 mole of CO_2 = 44g = 6.02×10^{23} molecules

This shows that:

44 g of CO_2 contain 6.02×10^{23} molecules of CO_2

1 g of CO_2 contain $\frac{6.02 \times 10^{23}}{44g}$ molecules

9 g of CO_2 will contain $\frac{6.02 \times 10^{23} \times 9g}{44g}$

$= 1.231 \times 10^{23}$ molecules of CO_2

By formula

$$\text{Number of molecules} = \frac{N_A \times \text{Mass of substance}}{\text{Formula mass}}$$

$$\text{Number of } (CO_2) \text{ molecules} = \frac{6.02 \times 10^{23} \times 9g}{44g}$$

$= 1.231 \times 10^{23}$ molecules of CO_2

Problem: 3.

Calculate the mass of one atom of carbon in grams

Solution: According to Avogadro's number
1 mole of C = 12g = 6.02×10^{23} atoms
this indicates that
6.02×10^{23} atoms of C weigh 12 g

$$1 \text{ atom of C} = \frac{12 \text{ g}}{6.02 \times 10^{23}} = \frac{12 \text{ g} \times 10^{-23}}{6.02}$$

$= 1.993 \times 10^{-23}$ g

Mass of one C-atom $= 1.993 \times 10^{-23}$ g

By formula

$$\text{Mass of one atom} = \frac{\text{Atomic mass in grams}}{N_A}$$

$$\text{Mass of one (C)atom} = \frac{12 \text{ g}}{6.02 \times 10^{23}} = 1.993 \times 10^{-23} \text{ g}$$

Problem: 4.

Calculate the mass of one molecule of water (H_2O) in gram.

Solution:

According to Avogadro's number

1 mole of H_2O = 18g = 6.02×10^{23} molecules

This indicates that

6.02×10^{23} molecules of (H_2O) weigh 18g

1 molecule of (H_2O) weighs $\frac{18\ g}{6.02 \times 10^{23}} = \frac{18\ g \times 10^{-23}}{6.02} = 2.90 \times 10^{-23} g$

By formula

$$\text{Mass of one molecule} = \frac{\text{Formula mass in grams}}{N_A}$$

$$\text{Mass of one } (H_2O) \text{ molecule} = \frac{18\ g}{6.02 \times 10^{23}} = 2.90 \times 10^{-23}\ g$$

Mass of one (H_2O) molecule in grams = 2.90×10^{-23} g

2.4 CHEMICAL REACTION OR CHEMICAL CHANGE

Any change, which alters the composition of a substances, is a chemical change. In this type of change one. or more new substances are formed from the original substances, for example, when iron (Fe) rusts, it reacts with oxygen (O) of air in presence of moisture to form red brown iron oxide (rust). Similarly, when coal burns, it forms smoke, gaseous products and ashes. The burning of coal is a chemical reaction (change)in which it combines with oxygen in air to form entirely new substances.

2.4.1 Types of Chemical Reactions:

Chemical reactions can be divided commonly into five different types.

1. Decomposition reactions.
2. Addition reaction (combination reaction).
3. Single displacement reaction.
4. Double displacement reaction.
5. Combustion reaction.

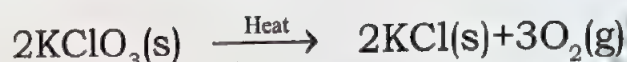
1. **Decomposition Reactions:**

A reaction in which a chemical substance breaks down to form two or more simpler substances is called a decomposition reaction. These reactions require some energy for decomposition.

For example: Calcium carbonate decomposes into calcium oxide and carbondioxide in presence of heat.

$$CaCO_3(s) \xrightarrow{Heat} CaO(s) + CO_2(g)$$

Similarly, potassium chlorate ($KClO_3$) on heating produces two simpler substances, potassium chloride and oxygen gas.

$$2KClO_3(s) \xrightarrow{Heat} 2KCl(s)+3O_2(g)$$

(2) Addition or Combination Reaction:

A reaction in which two or more substances combine to form a single substance is called an addition or combination reaction. These reactions are reverse of decomposition reactions.

For example: Calcium oxide (CaO) reacts with carbon dioxide (CO_2) to form calcium carbonate ($CaCO_3$).

$$CaO(s)+CO_2(g) \longrightarrow CaCO_3(s)$$

Another example is, when sodium reacts with chlorine gas it gives a new substance called common salt (NaCl).

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

(3) Single Replacement (Displacement) Reaction:

A reaction in which one atom or group of atoms of a compound is replaced by another atom or group of atoms is defined as displacement reaction. Some metals react with acids, bases or even water to displace hydrogen (H_2) gas.

For example: Zinc replaces hydrogen in hydrogen chloride (HCl) to give zinc chloride.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2(g)$$

Similarly sodium (Na) metal reacts with water to give sodium hydroxide and hydrogen gas.

$$2Na + 2(H - OH) \longrightarrow 2NaOH + H_2(g)$$

When chlorine reacts with a solution of potassium bromide, chlorine replaces bromine to form KCl and Br_2 vapours.

$$Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(g)$$

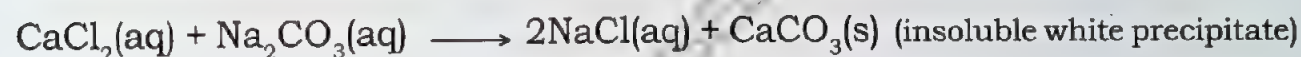
(4) Double Displacement Reaction:

It is a reaction in which two compounds exchange their partners, so that two new compounds are formed. In double displacement reaction usually there is an exchange of ionic radicals.

For example: When sodium chloride (NaCl) is reacted with silver nitrate ($AgNO_3$) solution, they exchange their partners to form two different compounds silver chloride (AgCl) and sodium nitrate ($NaNO_3$).

$$NaCl(aq) + AgNO_3(aq) \longrightarrow NaNO_3(aq) + AgCl(s) \text{ (insoluble white precipitate)}$$

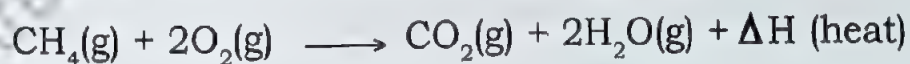
Consider another example, when, calcium chloride ($CaCl_2$) is reacted with sodium carbonate (Na_2CO_3) they exchange their partners to form two new compounds, sodium chloride and calcium carbonate ($CaCO_3$)(s).

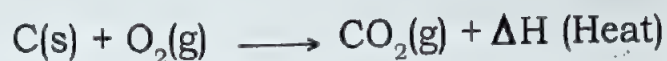
$$CaCl_2(aq) + Na_2CO_3(aq) \longrightarrow 2NaCl(aq) + CaCO_3(s) \text{ (insoluble white precipitate)}$$

Remember that neutrilization and hydrolysis reactions are also double displacement reactions. These reactions will be discussed in the 9th chapter on Acids, Bases and Salts.

(5) Combustion Reaction:

A reaction in which substances react with either free oxygen or oxygen of the air, with the rapid release of heat and flame, is called combustion reaction.

For example, when methane (CH_4), gas burns in air, it forms carbon dioxide gas (CO_2), water (H_2O) and heat.

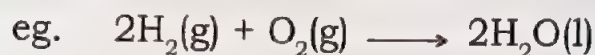
$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) + \Delta H \text{ (heat)}$$

Similarly, when carbon (C) burns in air, it produces carbon dioxide (CO_2) gas and heat

$$C(s) + O_2(g) \longrightarrow CO_2(g) + \Delta H \text{ (Heat)}$$

2.4.2 Chemical Equation:

Chemical equation is short hand method of describing (expressing) the chemical reaction, in terms of symbols and formulae of the substances involved in a chemical reaction.

The starting substances are called reactants and are always written on the left hand side, where as the substances which are produced (formed) are known as products and are always written on the right hand side of an equation. The reactants and products are separated from one another by using the single arrow (→) or double headed arrow (⇌), depending upon the kind (type) of reaction.

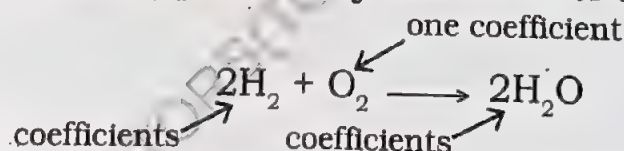
For illustration, when two molecules of hydrogen (H_2) combine (react) with one molecule of oxygen (O_2) to give two molecules of water (H_2O), instead of writing the full names of substances, chemist represents this chemical reaction in the form of following equation.

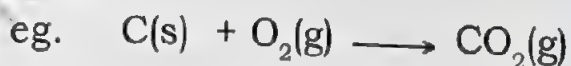
eg. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

Here, hydrogen and oxygen are called "Reactants" (substances, which are present before the chemical reaction) and water is the product (substances that result from the chemical reaction).

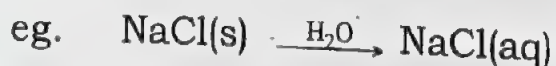
The numbers, in front of the formulas in a chemical equation are called co-efficients (they show the number of molecules that react with each other) where no co-efficient appears, only one number is considered.

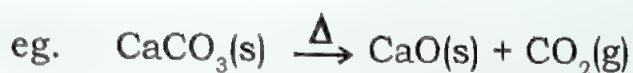
one coefficient

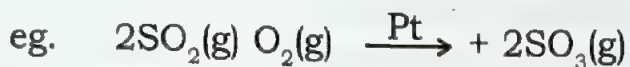
$2H_2 + O_2 \longrightarrow 2H_2O$

coefficients coefficients

The expressions (g), (l) and (s) are placed some times as subscript after the formulas of the reactants and products, indicating the state, gaseous, liquid and solid of reactants and products.

eg. $C(s) + O_2(g) \longrightarrow CO_2(g)$

Another expression, frequently used is (aq) for aqueous, showing that the substance is in the form of water solution.

eg. $NaCl(s) \xrightarrow{H_2O} NaCl(aq)$

Sometimes, reaction conditions are written over the arrow, for example, when reactants are heated, a capital Greek letter delta (Δ) may be placed over the arrow.

eg. $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

Similarly, if catalyst is used, this catalyst is shown over the arrow.

eg. $2SO_2(g)\ O_2(g) \xrightarrow{Pt} + 2SO_3(g)$

If the reactants are heated in the presence of catalyst, both the symbols are placed on the arrow.

eg. $2KClO_3(s) \xrightarrow{\Delta,\ MnO_2} 2KCl(s) + 3O_2(g)$

2.4.3 Writing of Chemical Equation:

Before we write a chemical equation we must know the composition of all reactants and products. Otherwise, we cannot introduce the proper formulae into equation. Knowing the formulae of all reactants and products, we balance the equation by knowing the fact that in a chemical reaction atoms are neither created nor destroyed (to conform with the law of conservation of mass). Thus the number of atoms of each element must be same on both sides of the arrow in a chemical equation.

Consider, when zinc (Zn) reacts with sulphuric acid (H_2SO_4) to form zinc sulphate $ZnSO_4$ and hydrogen (H_2) gas.

The chemical equation for this reaction can be represented as:

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Where the arrow ($\rightarrow$) is read as "gives" "produces", "yields" or "forms". The (+) sign on the left side of equation appears for "reacts with" and on the right side of an equation is read as "and". The reaction is assumed to proceed from left to right, as the arrow indicates.

2.4.4 The meaning of Chemical Equation:

The chemical equation gives the following important information about the chemical reactions.

i) The nature of reactants and products.
ii) The relative number of each i.e reactants and products.

2.4.5 Balancing of Chemical Equation:

All chemical equations must be balanced in order to comply with the law of conservation of mass. In balancing the chemical equation, to make the number of atoms of each element the same on both sides of the equation,

we can change the co-efficient (the number in front the formulas) but not subscript (the number within formulas). Remember that changing a subscript in a formula e.g. from H_2O to H_2O_2 change the identity of chemical compound. The substance hydrogen peroxide (H_2O_2) is quite different from water (H_2O). In contrast, changing a co-efficient, only changes the amount but not the identity of the substance. $2H_2O$ means two water molecules and $3H_2O$ means three water molecules and so on. Reduce the co-efficient to their smallest whole number values, if necessary, dividing them by common divisor. Thus the formulas of the reactants and the products in an equation remain same i.e. fixed and cannot be altered; so the only way of balancing an equation by taking appropriate numbers of molecules of the reactants and the products concerned.

Most chemical equations can be balanced by inspection method, that is, by trial and error method, with the experience, you should be able to balance any equation quickly. In general we can balance the equation by the following steps.

1. Write the correct formulae of all reactants on the left side and the formulae of products on the right side of an equation.
2. Balance the number of atoms on each side.
3. If the number of atoms may appear more on one side than the other, balance the equation by inspection method for this, multiply the formula by co-efficient so as to make the number of atoms, same on both sides of an equation.
4. The covalent molecules of hydrogen, oxygen, nitrogen and chlorine exist as di-atomic molecules eg. H_2, O_2, N_2 and Cl_2, rather than isolated atoms, hence we must write them as such in chemical equation.
5. Finally, check the balanced equation, to be sure that the number and kind of atoms are the same on both sides of equation.

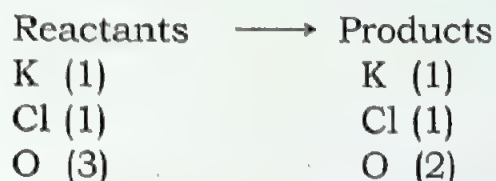
Let us consider a specific example. In the laboratory, oxygen (O_2) gas is conveniently prepared by heating potassium chlorate ($KClO_3$). The products are potassium chloride (KCl) and oxygen (O_2) gas.

To balance the equation.

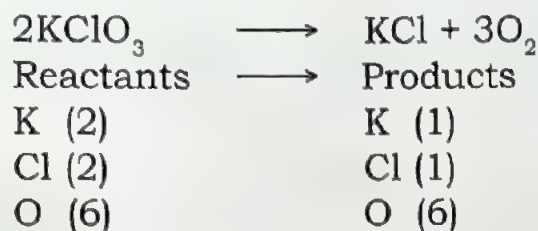
Write correct formulae of reactants on left side and formulae of the products on right side of an equation.

$$KClO_3(s) \longrightarrow KCl(s) + O_2(g)$$

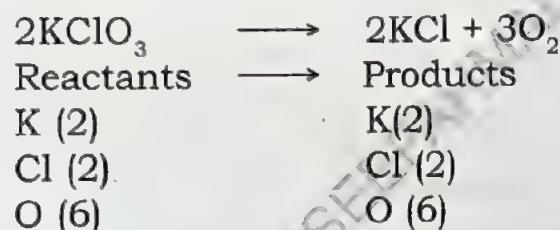
Balance the number of atoms on each side.

Reactants	$\longrightarrow$	Products
K (1)		K (1)
Cl (1)		Cl (1)
O (3)		O (2)

We see that (K) and (Cl) elements have the same number of atoms on both sides of the equation, but there are three oxygen atoms on the left and two oxygen atoms on the right side of the equation. We can balance the oxygen atoms by placing a co-efficient 2 in front of $KClO_3$ and 3 in front of O_2.

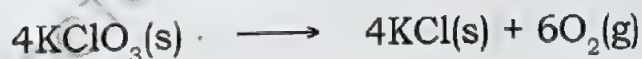
$$2KClO_3 \longrightarrow KCl + 3O_2$$

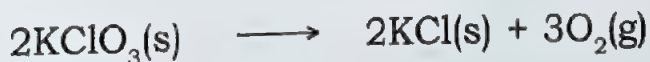
Reactants	$\longrightarrow$	Products
K (2)		K (1)
Cl (2)		Cl (1)
O (6)		O (6)

We balance the (K) and (Cl) atom by placing 2 in front of KCl.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

Reactants	$\longrightarrow$	Products
K (2)		K(2)
Cl (2)		Cl (2)
O (6)		O (6)

This equation is now balanced, because the number of atoms of each element are same on both sides of the equation.

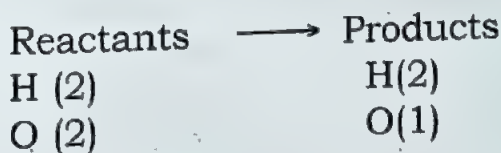
Note that this equation could also be balanced with co-efficients, that are multiple of 2 of each, for example:

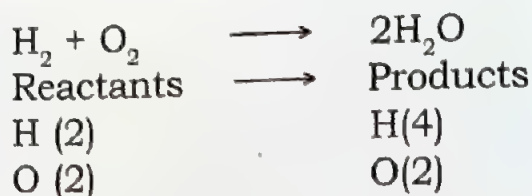
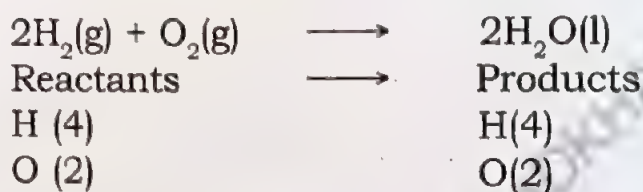
$$4KClO_3(s) \longrightarrow 4KCl(s) + 6O_2(g)$$

However, it is common practice to reduce the co-efficients to their smallest whole number value, here, this equation is divided by 2, so as to get smaller whole number ratio.

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

Consider, another example, when hydrogen burns in air (which contains oxygen) to form water, we can write the equation.

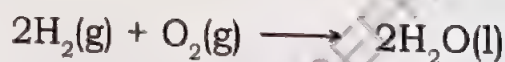
$$H_2(g) + O_2(g) \longrightarrow H_2O(l)$$

Balance the number of atoms on each side.

Reactants	$\longrightarrow$	Products
H (2)		H(2)
O (2)		O(1)

We see that hydrogen element has the same number of atoms on both sides of the equation, but there are two oxygen atoms on left side and one oxygen atom on the right side of the equation. We can balance the oxygen atoms by placing co-efficient 2, in front of water (H_2O).

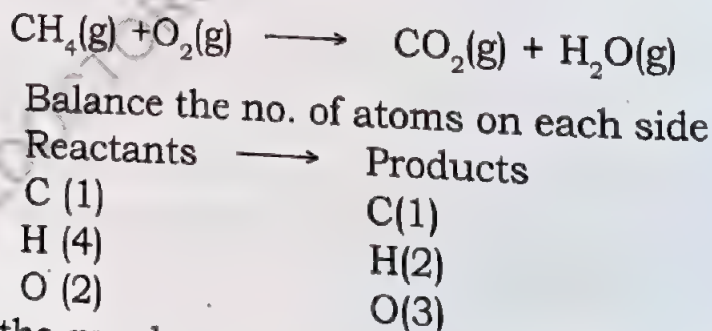
$H_2 + O_2$	$\longrightarrow$	$2H_2O$
Reactants	$\longrightarrow$	Products
H (2)		H(4)
O (2)		O(2)

Now we see that the number of oxygen atoms is same on both sides, but the number of hydrogen atoms can be balanced by placing 2 in front of (H_2) on left side of the equation, we get:

$2H_2(g) + O_2(g)$	$\longrightarrow$	$2H_2O(l)$
Reactants	$\longrightarrow$	Products
H (4)		H(4)
O (2)		O(2)

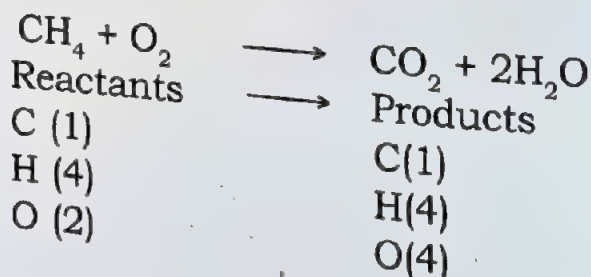
The equation is now balanced, because the number of atoms of each element are same on both sides of the equation.

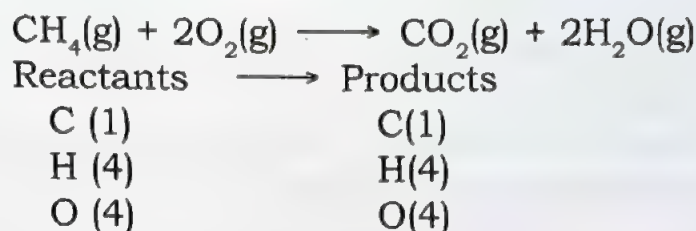
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

For further illustration, let us consider the combustion of natural gas methane (CH_4) in oxygen or air, which yields carbon dioxide (CO_2) and water (H_2O). We write from this information.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Balance the no. of atoms on each side

Reactants	$\longrightarrow$	Products
C (1)		C(1)
H (4)		H(2)
O (2)		O(3)

We see that the number of C-atoms is same on both sides but the number of hydrogen atoms on left side is (4), where as on right side is (2), we can balance it by placing a co-efficient 2 in front of water on right side, thus we get.

$CH_4 + O_2$	$\longrightarrow$	$CO_2 + 2H_2O$
Reactants	$\longrightarrow$	Products
C (1)		C(1)
H (4)		H(4)
O (2)		O(4)

Now, there is a difference in oxygen atoms, there are two oxygen atoms on left side but four oxygen atoms on right side, this can be balanced by placing co-efficient 2 in front of oxygen (O_2) on left side,

Thus we get

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

Reactants	$\longrightarrow$	Products
C (1)		C(1)
H (4)		H(4)
O (4)		O(4)

The equation is now balanced, because the number of atoms of each element are same on both sides of equation

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

2.4.6 Concept of Mole-Ratio Based on Balanced Chemical Equation:

As we have seen, the co-efficient in a chemical reaction represents the number of moles (molecules) and not masses of molecules. However in chemical reaction, the amount of the reactants needed can not be determined by counting molecules directly. Counting is always done by weighing. To find out the masses of reactants and products, in a balanced equation, the first step is to find out the mole-ratios in the balanced equation. Then convert the moles of reactants or products into mass.

The following steps will help in calculating the amount of reactant or products.

(1) Balance the equation for the given chemical reaction.

(2) Use the co-efficients in the balanced equation to get the mole ratio.

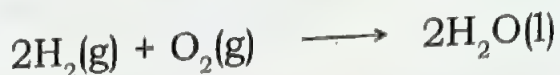
(3) Use the mole-ratio to calculate the number of moles of desired reactants or products.

(4) Convert the moles of reactants or products into mass, if required by the problem.

The following example illustrates the use of the above four steps, in solving the mole-ratio problems.

Calculating The Amount of Reactants.

Hydrogen (H_2) reacts with oxygen (O_2) to form water. The equation for this reaction is:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

i) How many moles of oxygen are needed to react with 4.5 moles of hydrogen?

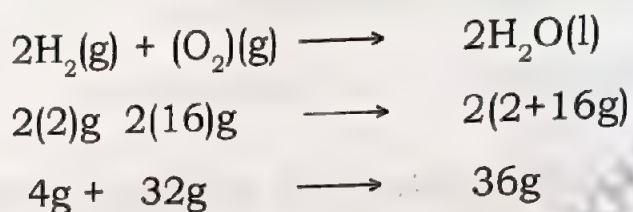
ii) How many grams of hydrogen will completely react with 100g of oxygen to form water?

Answer.

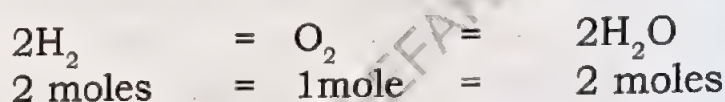
Step-(1) Balance the equation for chemical reaction.

$$2H_2(g) + (O_2)(g) \longrightarrow 2H_2O(l)$$

$$2(2)g \quad 2(16)g \longrightarrow 2(2+16g)$$

$$4g + 32g \longrightarrow 36g$$

Step-(2) Set up the mole ratio, since 2 moles of hydrogen react with 1 mole of oxygen to produce 2 moles of water.

$$2H_2 = O_2 = 2H_2O$$

$$2 \text{ moles} = 1\text{mole} = 2 \text{ moles}$$

Step-(3) (i). Calculate the moles of oxygen as follows.

The balanced eq. shows that:

2 moles of hydrogen require 1 mole of oxygen

1 mole of hydrogen requires $\frac{1}{2}$ mole of oxygen

4.5 moles of hydrogen require $\frac{1 \times 4.5}{2} = 2.25$ moles

$= 2.25$ moles of oxygen

ii) 32g of oxygen react with 4g of hydrogen

1g of oxygen reacts with $\frac{4g}{32}$

100g of oxygen react with $\frac{4 \times 100}{32} = 12.5g$

Result: i) No. of moles of oxygen = 2.25 moles.
ii) 100 g of oxygen require = 12.5g of hydrogen.

Calculating the Amount of Products:

Methane (CH_4) is a common fuel used for cooking. The chemical reaction is,

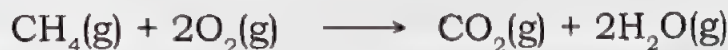
$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

i) How many moles of CO_2 will be produced by complete combustion of 10 moles of CH_4.

ii) What mass of H_2O vapours will be formed?

Answer.

Step (1) Balance the equation for chemical reaction

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

Step (2) Set up the mole ratio, the balance equation shows that 1 mole of methane (CH_4) reacts with 2 moles of oxygen (O_2) to produce 1 mole of carbon dioxide (CO_2) and 2 moles of water (H_2O)

$$\underset{1\text{mole}}{CH_4} + \underset{2\text{moles}}{2O_2} \longrightarrow \underset{1\text{mole}}{CO_2} + \underset{2\text{moles}}{2H_2O}$$

Step (3)

i) Calculate the moles of (CO_2) as follows.
The balanced equation shows that:
1 mole of CH_4 on combustion produces 1 mole of CO_2
10 moles of CH_4 on combustion produces 1x10 = 10 moles of CO_2

ii) Mass of H_2O
The balanced equation indicates that:
1 mole of CH_4 on combustion produce 2 moles of H_2O
10 moles of CH_4 on combustion produces 2x10=20 moles of H_2O
Mass of H_2O = no. of moles x m. mass
Mass of H_2O = ??
No. of moles = 20 moles
M. mass of H_2O = 2+16 = 18g
$\therefore$ Mass of H_2O = 20 x 18g = 360g

Result. No. of moles of CO_2 = 10 moles
Mass of H_2O formed = 360 grams

SUMMARY

1. There are several laws that govern the composition of matter and chemical reaction. These are, the law of conservation of mass, the law of definite proportions, the law of multiple proportions and the law of reciprocal proportions.
2. Atomic mass is the average mass of naturally occurring isotopes, which is compared to the mass of one atom of carbon-12 a.m.u.
3. The molecular mass of a molecule is the sum of the atomic masses of all the atoms present in that molecule.
4. Formula mass of the substance is sum of the atomic masses of all the atoms present in the formula unit of the substance.
5. Molar mass is the mass in grams of one mole of a substance.
6. Mole is the amount of substance that contains as many elementary particles (atoms, ions, molecules or formula units) as there are atoms in 12g of carbon-12 isotope.
7. Avogadro's number, the number of particles i.e. atoms, ions, molecules or formula units in one mole of substance equal to $6.02x10^{23}$.
8. The empirical formula is the formula which describes the smallest or the least ratio of the combining atoms of different elements present in a molecule.
9. The Molecular formula shows the actual number of the combining atoms of all the elements present in a molecule and is always an integer multiple of empirical formula.
10. Chemical reaction is a change in which the composition of substance is changed and new substances are formed. Chemical reactions can be divided into five different types. Decomposition reaction, combination reaction, single replacement reaction, double replacement reactions and the combustion reactions.
11. A chemical equation represents the chemical change or reaction, showing reactants on the left side of an arrow and products on the right. The equation is balanced by placing co-efficients in front of the formulas, so the number of atoms of each element are same on each side of an arrow.
12. Amounts of reactants and products formed can be calculated from the balanced equation for a reaction by using the mole-ratios relating the reactants and products.

EXERCISE

1. Fill in the blanks :

(i) 18 grams of H_2O contains .... molecules.

(ii) A change which alters the composition of a substance is called.......

(iii) A reaction in which a chemical substance breaks down to form two or more simpler substances, is called................

(iv) The reaction of NaCI with $AgNO_3$ is given as

$$NaCI(aq) + AgNO_3(aq) \longrightarrow AgCI(S) + NaNO_3(aq)$$

is the reaction of the type....................

(v) When metals react with acids or water then produce .... gas.

(vi) .... is the reaction in which two or more substances combine together to form a single substance.

(vii) A reaction in which a substance burns in oxygen to produce heat and flame is called.........

(viii) .... is the short hand method to describing a chemical reaction.

(ix) The reaciton $Zn + 2HCI \longrightarrow ZnCI_2 + H_2$ (g) is the .... replacement reaction.

2. Tick the correct answer :

(i) Mass is neither created nor destroyed during a chemical change, is the statement of:

a) Law of conservation of mass.
b) Law of definite proportions.
c) Law of multiple proportions.
d) Law of reciprocal proportions.

(ii) A given compound always contains exactly the same proportion of elements, by mass, is the statement of:

a) Law of conservation of mass.
b) Law of definite proportions.
c) Law of multiple proportions.
d) Law of reciprocal proportions.

(iii) The average mass of natural mixture of isotopes, which is compared to the mass of one atom of C-12 a.m.u, is called:

a) Atomic number. b) Mass number.
c) Atomic mass. d) None of these.

(iv) A formula that gives only the relative number of each type of atoms in a molecule, is called:
a) Empirical formula.
b) Molecular formula.
c) Molecular mass.
d) Formula mass.

(v) A formula that indicates actual number and type of atoms in a molecule, is called:
a) Empirical formula.
b) Molecular formula.
c) Molecular mass.
d) Formula mass.

(vi) The sum of the atomic masses of all atoms in a molecule is called:
a) Empirical formula.
b) Molecular formula.
c) Molecular mass.
d) Formula mass.

(vii) The sum of the atomic masses of all atoms in a formula unit of substance is called:
a) Empirical formula.
b) Molecular formula.
c) Molecular mass.
d) Formula mass.

(viii) The mass of (1) mole of substance expressed in grams, is called:
a) Empirical formula.
b) Molecular formula.
c) Molecular mass.
d) Molar mass.

(ix) 44 a.m.u. of CO_2 is equal to:
a) Molar mass.
b) Atomic mass.
c) Molecular mass.
d) Mass number.

(x) 5 moles of H_2O are equal to:
a) 80g.
b) 90g.
c) 100g.
d) 90 a.m.u.

3. Write answer of the following questions :

(i) State the law of conservation of mass? Describe Landolt experiment?
(ii) State the law of definite proportions in your own words.
(iii) What is law of multiple proportions? Explain with examples.
(iv) State the law of reciprocal proportion and illustrate it with examples.
(v) What is empirical formula? Give an example?
(vi) What is molecular formula? Give an example?

(vii) Can one substance have the same empirical formula and molecular formula? Explain with examples.

(viii) What is the difference between empirical formula and molecular formula?

(ix) What is atomic mass unit?

(x) The value of atomic mass of carbon in periodic table is 12.011 a.m.u, rather than 12.00 a.m.u? Explain?

(xi) The atomic masses of ${}_{17}Cl^{35}$ (75%) and ${}_{17}Cl^{37}$ (25%). Calculate the average atomic mass of chlorine.

(xii) How many atoms are there in 5 moles of sulphur?

(xiii) What is a mole? What is the molar mass of substance? Find out the molar mass of SO_2?

(xiv) What is Avogadro's number? What number of oxygen atoms are present in 4g of oxygen?

(xv) What does Avogadro's number represent?

(xvi) What is mass in grams of a single atom of each of the following elements?

(a) Carbon(C) (b) Magnesium (Mg) (c) Calcium (Ca).

(xvii) What is mass in grams of 1×10^{20} atoms of Na?

(xviii) Define the following.

(a) Molecular formula mass

(b) Formula mass

(c) Molar mass

(xix) Calculate the molecular mass (in a.m.u) of each of the following substances.

(a) H_2O (b) H_2O_2 (c) C_6H_6 (d) C_2H_6O.

(xx) Calculate the formula mass (in a.m.u.) of each of the following:

(a) KNO_3 (b) Al_2O_3 (c) $CaCO_3$ (d) $MgCl_2$.

(xxi) Calculate the molar mass of the following substances.

(a) S_8 (b) CS_2 (c) $CHCl_3$ (Chloroform).

(d) CH_3-COOH (Acetic acid).

(xxii) The formula for rust is Fe_2O_3. How many moles of (Fe) are present in 30g of rust?

(xxiii) Define the following terms:

(a) Chemical reaction.

(b) Reactants.

(c) Products.

(xxiv) What is chemical equation? What is a co-efficient? Give an example of balanced equation.

(xxv) What is combination reaction? Give an example.

(xxvi) What is decomposition reaction? Will two or more elements always be the products of this type of reaction? Explain with examples.

(xxvii) What is single replacement reaction? Give an example?

(xxviii) Explain double displacement reaction with examples?

(xxix) Balance the following equations by inspection method?

(a) $C + O_2 \longrightarrow CO$

(b) $CO + O_2 \longrightarrow CO_2$

(c) $KNO_3 \longrightarrow KNO_2 + O_2$

(d) $NaH\ CO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$

(e) $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2O + CO_2$

(f) $NH_3 + O_2 \longrightarrow NO + H_2O$

(xxx) Which of the following reaction is either a decomposition reaction or combination reaction?

(a) $MgCO_3 \longrightarrow MgO + CO_2$

(b) $C_2H_4 + H_2 \longrightarrow C_2H_6$

(c) $BaCO_3 \longrightarrow BaO + CO_2$

(d) $N_2 + 3H_2 \longrightarrow 2NH_3$

(xxxi) Balance the equation and decide which one is single replacement reaction.

(a) $C_2\ H_2 + H_2 \longrightarrow C_4H_8$

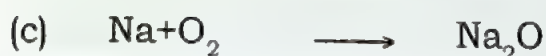
(b) $Ca + H_2O \longrightarrow Ca\ (OH)_2 + H_2$

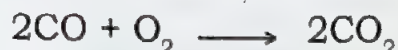
(c) $C_2H_5\text{-}OH + Na \longrightarrow C_2H_5ONa + H_2$

(xxxii) Balance the following equation, which of them is a decomposition reaction or combination reaction.

(a) $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

(b) $NO_2 \longrightarrow NO + O_2$

(c) $Na + O_2 \longrightarrow Na_2O$

(xxxiii) Consider the combination of (CO) with oxygen (O_2) gas.

$$2CO + O_2 \longrightarrow 2CO_2$$

Calculate the number of moles of (CO_2) produced when 50 moles of oxygen are reacted with all of (CO)?

(xxxiv) Calcium carbonate ($CaCO_3$) on heating gives calcium oxide (CaO) and CO_2 gas.

$$Ca\,CO_3 \longrightarrow CaO + CO_2$$

Calculate how many grams of calcium oxide (CaO) can be obtained by heating 8 moles of $CaCO_3$?

(xxxv) Silicontetrachloride ($SiCl_4$) can be prepared by heating (Si) in chlorine gas (Cl_2).

$$Si + 2Cl_2 \longrightarrow SiCl_4$$

If we want to prepare 10 moles of ($SiCl_4$) how many moles of molecular chlorine (Cl_2) will be used in the reaction.

(xxxvi) Fermentation is chemical decomposition, in which glucose ($C_6H_{12}O_6$) is converted into ethyl alcohol (C_2H_5-OH) and carbon dioxide (CO_2).

$$(C_6H_{12}O_6) \xrightarrow{\text{Fermentation}} 2C_2H_5\text{-}OH + 2CO_2$$

What will be the amount of ethyl alcohol in grams and moles, which can be obtained by fermentation of 5000g of glucose.

CHAPTER-3

ATOMIC STRUCTURE

You will learn in this chapter about:

* Dalton's atomic theory.
* Modern atomic theory.
* Fundamental particles of atoms.
* Discovery of Cathode particles i.e electrons.
* Discovery of Proton.
* Discovery of Neutron.
* Radioactivity.
* Rutherford's Atomic Model.
* Bohr's atomic Model.
* Atomic Number and Mass Number.
* Isotopes.
* Electronic Configuration based on Bohr's Model.
* Electronic Configuration.

3.1 INTRODUCTION

In the fifth century (B.C.) the Greek Philosopher Democritus expressed the belief that all matter consists of very small indivisible particles, which he named **ATOMOS** (Greek =Atomos = not cutable = indivisible) nowadays called atoms.

3.1.1 Daltons Atomic Theory:

In (1808) John Dalton, an English school teacher and Chemist suggested the fundamental ideas of atomic theory, which explains the chemical nature of matter and the existence of atoms. It is known as Dalton's atomic theory. The important postulates are as following:

1. All elements are made up of small indivisible, indestructible particles called atoms.
2. All atoms of a given element, are identical in all respects, having same size, mass and chemical properties. But the atoms of one element differ from the atoms of other element.
3. Compounds are formed when atoms of more than one element combine in a simple whole number ratio.
4. A chemical reaction is a rearrangement of atoms, but atoms themselves are not changed, this means that atoms are neither created nor destroyed in chemical reactions.

3.1.2 Modern Atomic Theory:

Daltons atomic theory assumed that atoms of elements are indivisible and that no particles smaller than atoms existed. But as the time passed new experimental facts led to the modification and extension of Dalton's atomic theory. Atom is a complex organisation, composed of even smaller particles called sub-atomic particles (fundamental particles). These are electrons, protons and neutrons. Dalton's view that all atoms of an element have the same mass is modified in the light of discovery of isotopes. Even then, we can say that the Dalton's atomic theory was largely successful in explaining the laws of chemical combinations.

3.1.3 Fundamental Particles of an Atom:

The atom was generally identified as the *smallest particle* of an element, consisting of sub atomic particles, the electrons, protons and neutrons. The first hint about the sub-atomic particles, came with the discovery of electron by M. Faraday (1832), William Crooks (1879) and J.J. Thomson (1897). Later, the second sub-atomic particle, the proton was identified and isolated by Goldstein, German scientist (1886) and Ernest Rutherford (1919). Finally an English scientist James Chadwick revealed the third particle the neutron in 1932. The structure of atom as we know it today, is because of these findings.

3.2. DISCOVERY OF ELECTRONS

The fundamental particle carrying a negative charge was discovered in 1897, by the British physicist J.J. Thomson. The apparatus used for this type of experiment is called the discharge tube (Neon sign and T.V. tube are examples of discharge tube) which consists of a glass tube, fitted with two metal electrodes connected to a high voltage source and a vaccum pump.

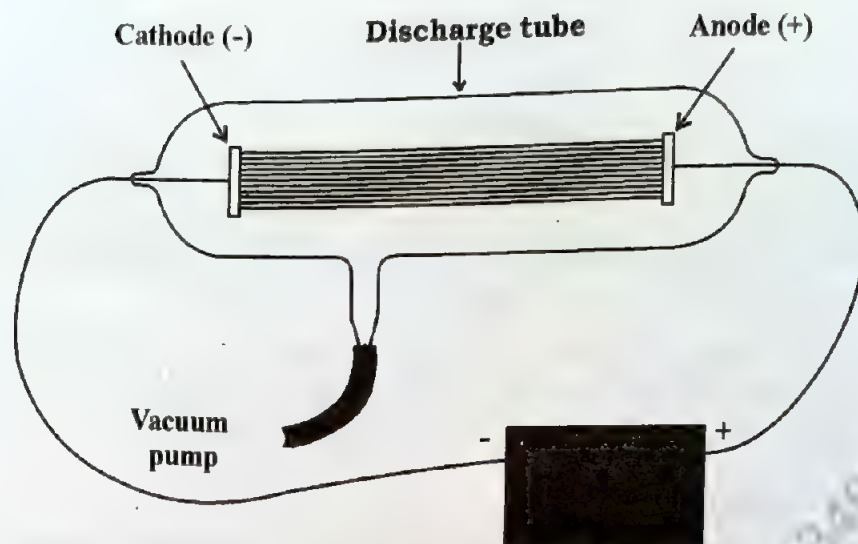



Fig: 3.1 A simple gas discharge tube

When the tube is evacuated and a current of high potential is passed between the electrodes, streaks of bluish light extending from negative electrode (cathode) towards positive electrode (anode). The rays appear to travel in straight lines, from the cathode to anode, cause the wall at the opposite end of the tube glow where they strike. These rays were called cathode rays. Thomson showed that these rays were deflected towards the positive plate in electric and magnetic field. This shows that these rays consisted of negatively charged particles. The name electron was given to these units of negative charges. Electrons were obtained irrespective of the nature of cathode or the gas in the discharge tube. This proves that electrons are constituents of all matter.

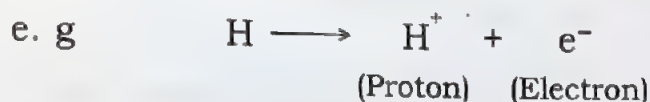
Properties of Cathode Rays:

1. They cast shadows of objects placed in their path towards the anode, proving that they travel in straight lines.
2. They *cause a light paddle* wheel to rotate (revolve). Showing that they are material particles.
3. These rays although invisible cause some material to glow or produce fluorescence.
4. They are deflected towards the positive plate in an electric field, showing that they are negatively charged particles.
5. The $\left(\frac{\text{charge}}{\text{mass}}\right)$ e/m ratio of cathode particles is 1.7588×10^{8} c/g (coulomb per gram), i.e same for all electrons, regardless of any gas in the tube.

6. They can produce mechanical pressure indicating they possess kinetic energy (K.E.).

3.2.1. Discovery of Protons:

Since atoms are electrically neutral and electrons carry negative charge, it follows that for each electron, there must be one equivalent positive charge to neutrilize that electron. This particle is called a proton. It is one of the fundamental units of structure of all atoms. The simplest atom of hydrogen (H) is therefore made up of one electron and one proton.

e. g $\quad H \longrightarrow \underset{\text{(Proton)}}{H^+} + \underset{\text{(Electron)}}{e^-}$

Protons were first observed in apparatus similar to cathode rays tube, with a perforated cathode by German Physicist Goldstein in 1886,their existence was verified and their properties were investigated in 1897 by J.J. Thomson.

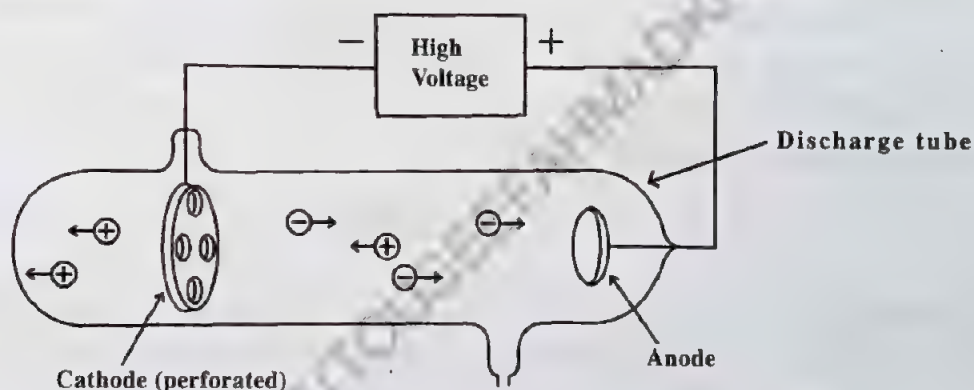



Fig: 3.2 Goldstein's apparatus for studying positive particles

Properties of positive rays :

1. These rays also travel in straight line from anode to cathode.
2. These are deflected towards the negative plate when passed through an electric field, showing that these carry a positive charge.
3. The $\frac{\text{charge}}{\text{mass}}$ i.e., e/m ratio of positive particle is much smaller than that for electron and it varies with the nature of gas in the tube.

3.2.2 Discovery of Neutrons:

In 1932 the English Physicist James Chadwick discovered a third type of fundamental particle of atomic structure through artificial radioactivity, further it will be discussed in the definition of artificial radioactivity.

Properties of Electrons Protons and Neutrons:

Electron :

Electron is negatively charged particle. The charge measured in coulombs. The electric charge is equal to 1.602×10^{-19} coulombs. The electron carries a negligible mass. Its mass is $\frac{1}{1836}$ part of that of proton. The actual mass of an electron is 9.109×10^{-31} Kg.

Proton:

Proton is positively charged particle. The actual charge on a proton is 1.602×10^{-19} coulombs. A proton is 1836 times heavier than electron. The actual mass of a proton is 1.672×10^{-27} Kg.

Neutron:

The neutron is neutral particle, i.e. it has no charge. Its mass is almost equal to that of proton. The actual mass of a neutron is 1.76×10^{-27} Kg. Properties of three sub-atomic particles of the atoms are summarized in table No. (3.1)

Table 3.1

PARTICLE	MASS			CHARGE		
	GRAMS	Kilo GRAMS	a.m.u.	COULOMBS	e.s.u.	CHARGE UNIT
Electron	9.109390×10^{-28}	9.109390×10^{-31}	0.0005485	$-1.602177 \times 10^{-19}$	-4.8×10^{-10}	– 1
Proton	1.672623×10^{-24}	1.672623×10^{-27}	1.007276	1.602177×10^{-19}	4.8×10^{-10}	+ 1
Neutron	1.764929×10^{-24}	1.764929×10^{-27}	1.008664	NONE (0)	NONE (0)	NONE(0)

3.2.3 Radioactivity:

The first conclusive evidence that atoms are complex rather than **indivisible** as stated in the atomic theory, came with the discovery of radioactivity by Henry Becquerel, a French Physicist in 1896.

Definition of Radioactivity:

Radioactivity is the spontaneous disintegration of nucleus of an atom, in which invisible radiations are emitted from the nucleus of atoms. The substances which emit such kind of radiations are known as radio-active

elements and the phenomenon is termed as Radioactivity.

Nature of Radioactivity: (Types of rays)

The British physicist Ernest Rutherford in 1902 determined the nature of Radioactive rays by the following experiment and showed that, it is composed of three types of rays.

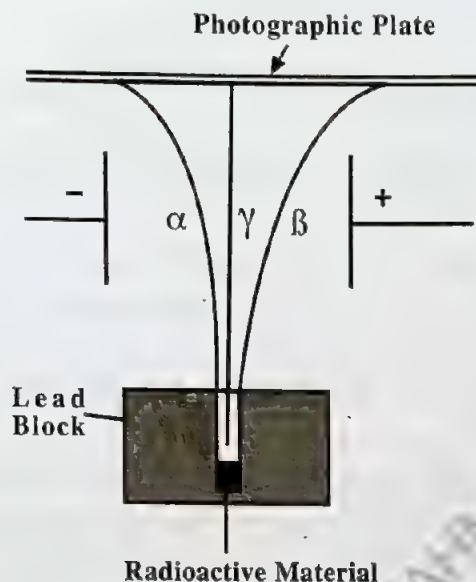



Fig.3.3 Shows separation of alpha, beta and gamma rays.

A sample of radioactive substance was placed in a lead block, between the two oppositely charged plates (electric field). The radiaitons were resolved into three components. One component was deflected towards the negative plate, proving that it carried a positive charge. These were named α-rays. The second component deflected towards the positive plate, showing that it carried a negative charge. These were named β-rays. The third type carried no charge no mass and were not deflected in the electric field. These were, named γ-rays.

3.3 RUTHERFORD ATOMIC MODEL

Lord Rutherford in 1911, carried out series of experiments. He passed a beam of α-particles through a very thin gold metal foil. He found that most of the α-particles passed through it without any deflection. However some of them deflected at large angles and very few of them bounced back.

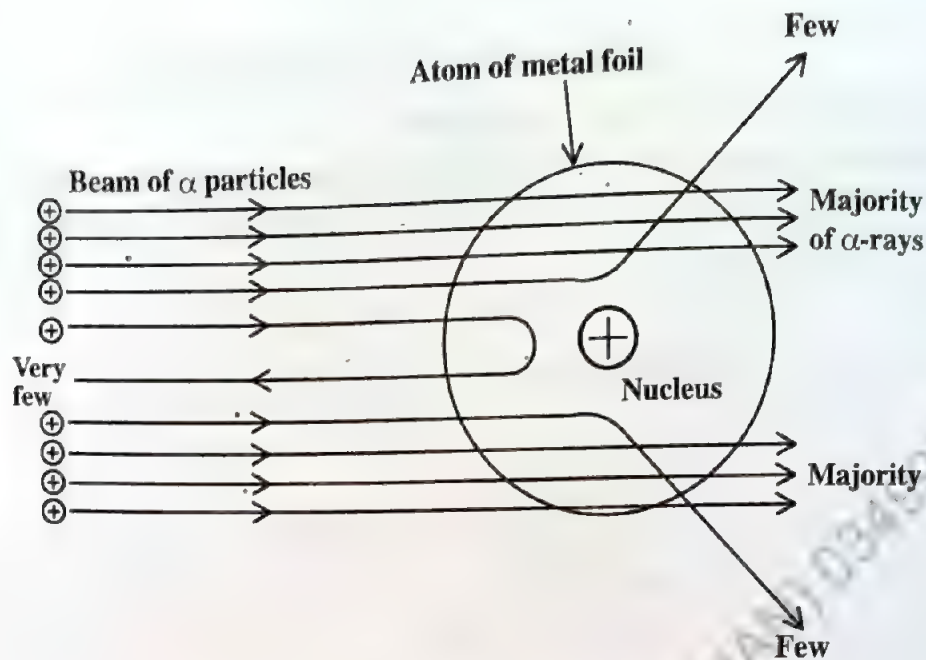



Fig: 3.4 Rutherford experiment of gold metal foil.

From the result of his experiment, Rutherford concluded that:

(1) The volume occupied by an atom must be largely empty as most of the ∝-particles passed through the foil un-deflected.

(2) The positive charge, in the atom is concentrated in extremely dense region which he called the nucleus. This was from the fact that ∝-particles after collision with a heavy positively charged nucleus had bounced back.

From the above observations, Rutherford proposed that the atom consists of very small, positively charged nucleus in which the most of the mass of the atom is concentrated. The rest of the volume is empty space, However this space is not completely empty and that in it electrons revolve around the nucleus.

According to the Rutherford model, an atom consisted of two parts.

1. Nucleus.
2. Extra nuclear part.

The proton and neutron reside in the nucleus. Since the protons are positive charged particles, therefore, the nucleus has positive charge. Further since the weight of the atom due to presence of protons and neutrons, as these particles are residing in the nucleus, the weight of the atom is concentrated in the nucleus.

The electrons are revolving around the nucleus in the extra nuclear part in various orbits, which are also called as shells, or energy levels.

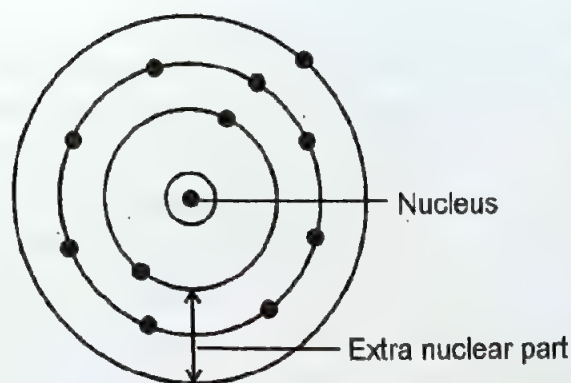



Fig: of Atom.

3.3.1 Weaknesses or Defects in Rutherford's Atomic Model:

1. According to the classical physics, since electron is revolving around the nucleus constantly it should lose energy and it ultimately falls into the nucleus.
2. If the revolving electron emits energy continuously, then there would be a continuous spectrum, but in contrast to it, we get line spectrum from the atoms of element.

3.3.2 Neil Bohr's Atomic Model:

After Planks and Einstein's discoveries, Niel Bohr, a Danish Physicist in 1913 offered a theoretical explanation of line spectra. The important assumptions for the atomic structure are given below:

1. Neil Bohr adopted Planks idea, that energies are quantized. He proposed that the electrons in atoms move only in certain allowed energy levels (energy states), so an electron in an allowed energy state will not radiate energy continuously and therefore will not fall in the nucleus.
2. That the atom radiates energy as a light only when the electron jumps from higher energy level (E_2) to lower energy level (E_1). The quantity of energy radiated is in discrete quantity, called quanta. A quantum of energy is directly proportional to the frequency of the radiation.

i.e. $\Delta E = E_2 - E_1 = h\nu$

Where h = Planks constant

ν = is the frequency of the radiation.

3.4 ATOMIC NUMBER (Z) AND MASS NUMBER (A)

Why the atoms of one element differ from the atoms of another

element? For example, how does an atom of carbon (C) differs from an atom of nitrogen (N). All the atoms can be identified by the number of protons they contain. No two elements have the same number of protons. *Thus the number of protons in the nucleus of an atom is called the atomic number.* In neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates, the *total number of electrons outside the nucleus. The atomic number is generally denoted by (Z).*

For example, atomic number of carbon (C) is 6. This means that each carbon atom has 6 protons and 6 electrons in it. The atomic number of nitrogen (N) is 7, it means that nitrogen atom has 7 protons and 7 electrons in it.

Atomic number = Z = number of protons in the nucleus or total number of electrons around the nucleus.

Atomic number (Z)is written as subscript on the left hand side of the chemical symbol e.g $_6C$ and $_7N$

3.4.1 Mass Number (A):

The nuclei of atoms contain both protons and neutrons, except the ordinary hydrogen atom, which consists of a single proton. The total sum of the protons and neutrons in the nucleus of an atom is called the mass number and is denoted by A. For example, the sodium (Na) atom has atomic number 11 and the mass number 23. It indicates that sodium atom has 11 protons and 12 neutrons. The mass number "A" is written as superscript on the left hand side of the chemical symbol.

e.g $^{23}_{11}Na$

Mass number (A) = Number of Protons (Z)+ Number of neutrons (N).
or mass no. A = (Z+N)
and number of neutrons = (A-Z)

3.4.2 Isotopes:

Atoms of the same element having the same atomic number but different atomic masses are called isotopes.

For example, chlorine element is composed of two kinds of chlorine atoms. These have masses 35 and 37. Both type of chlorine (Cl) atoms have atomic number 17 which means that each chlorine atom has 17 protons in the nucleus. The difference is only in the number of neutrons. Thus Cl-35 has 18 neutrons and Cl-37 has 20 neutrons. They are denoted by writing the atomic number as a sub-script and mass number as a superscript on the left hand side of chemical symbol..

e. g. $^{35}_{17}Cl$ and $^{37}_{17}Cl$

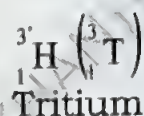
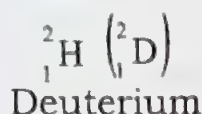
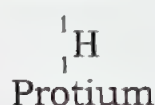
It is important to remember that the different isotopes of the same element differ only in the number of neutrons in the nucleus. Since chemical properties of an element depends on the electrons in the shells, the isotopes of an element have the same chemical properties but they differ in physical properties.

Isotopes of Hydrogen:

There are three isotopes of hydrogen. These are known as protium, deuterium and tritium. Protium has one proton and no neutron in the nucleus. Deuterium has one proton and one neutron in the nucleus. Tritium has one proton and two neutrons in the nucleus.

$^{1}_{1}H$ Protium

$^{2}_{1}H$ $(^{2}_{1}D)$ Deuterium

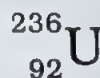
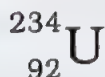
$^{3}_{1}H$ $(^{3}_{1}T)$ Tritium

Isotopes of Oxygen:

Oxygen atom has also three isotopes having mass number 16, 17, and 18. It shows that these atoms contain 8,9 and 10 neutrons respectively. Thus there are three isotopes of oxygen, i.e.

$^{16}_{8}O$ $^{17}_{8}O$ $^{18}_{8}O$

Isotopes of Uranium: (U)

There are three common isotopes of uranium with mass numbers 234, 235 and 236, respectively.

$^{234}_{92}U$ $^{235}_{92}U$ $^{236}_{92}U$

Some of the isotopes of C:

Symbol	No. of protons	No. of electrons	No. of neutrons
$^{12}_{6}C$	6	6	6
$^{13}_{6}C$	6	6	7
$^{14}_{6}C$	6	6	8

3.4.3 Applications of Isotopes:

Isotopes are used as tracers in physical, chemical and biological researches as well as in the treatments and dignoses of disease like cancer. Since radio isotopes are easily detected, it is relatively easy to trace their movements, even through it is a complicated system.

3.5. ELECTRONIC CONFIGURATION BASED ON BOHR'S MODEL

When atoms react, it is actually the electrons that interact. For this reason, the arrangement of electrons is responsible for the structure of atom. The arrangement of electrons refers not only to the number of electrons that an atom possesses, but also their distribution around the nucleus.

According to Bohr's theory, the electron move in a specific circular orbits around the nucleus, much like the earth moves around the sun. These orbits are called energy levels or shells and are designated as K, L, M, N, O, P orbits and represented by 1, 2, 3, 4, 5, 6, starting from one nearest to the nucleus.

Therefore, K shell is the first shell, L shell is second shell, M shell is the third shell, and so on. The maximum number of electrons in the given shell is governed by the formula $2n^2$, where n is the number of orbits (shell or energy level). Thus first orbit can have a maximum of two electrons (n=1). The second orbit can have a maximum of eight electrons (n=2) and the third orbit can have a maximum of eighteen electrons (n=3) and so on.

The arrangement of electron in the first twenty elements in the periodic table is as under :

Table 3.3 Shows the Arrangement of Electrons in the First 20 Elements.

Elements	Atomic Numbers	Electrons in the Shells			
		K or 1st	L or 2nd	M or 3rd	N or 4th
H	1	1			
He	2	2	–	–	–
Li	3	2	1		
Be	4	2	2	–	–
B	5	2	3	–	–
C	6	2	4	–	–
N	7	2	5	–	–
O	8	2	6	–	–
F	9	2	7	–	–
Ne	10	2	8	–	–

Na	11	2	8	1	–
Mg	12	2	8	2	–
Al	13	2	8	3	–
Si	14	2	8	4	–
P	15	2	8	5	–
S	16	2	8	6	–
Cl	17	2	8	7	–
Ar	18	2	8	8	–
K	19	2	8	8	1
Ca	20	2	8	8	2

Summary

1. Dalton in 1808 published his work under the name Dalton's Atomic Theory.
2. The fundamental particles in an atom are electrons, protons and neutrons. Electrons are negatively charged particles, protons are positively charged particles and neutrons are neutral particles.
3. Protons and neutrons are found in the nucleus of an atom. Electrons move around the nucleus in regions called shells or energy levels.
4. The atomic number of an element is the number of protons in the nucleus of an atom, it determines the identity of an element.
5. The mass number is the sum of the number of protons and the number of neutrons in the nucleus.
6. Isotopes are atoms of same element that have the same atomic number (number of protons) but different number of neutrons.
7. Substances which give off invisible rays that affect photographic plates in the same way as light does, are called radio-active substances and the phenomenon is termed as Radio-activity. The radiation from Radioactive substances are of three types (a) Alpha particles or helium nuclei (b) Beta particles or electrons and (c) gamma rays or high energy x-rays.
8. According to Rutherford atomic model, an atom consists of small, dense, and positively charged centre, called nucleus, which is surrounded by electrons at relatively greater distance from it.

9. Neil Bohr proposed a model of the atom in which the electrons in a hydrogen atom moves around the nucleus in fixed orbits. The line spectrum is only produced when an electron jumps from higher orbit to lower orbit.
10. An electronic configuration is a particular distribution of electrons among the different shells of an atom.

EXERCISE

1. Fill in the blanks :

(i) .... model, says that atom consists of small, dense, positively charged nucleus which is surrounded by electrons, revolving arround it.

(ii) Atomic number of sodium is...................

(iii) Number of proton + number of neutrons is the .... of an element.

(iv) .... are the atoms of the same elements, having same number of protons but different number of neutrons.

(v) The number of isotopes of hydrogen is.............

(vi) .... is the number of positive charges in the nucleus of an atom.

(vii) A-Z indicates the number of .... in the nucleus of an atom.

(viii) Z = number of protons in the nucleus of an atom = number of .... in a neutral atom.

2. Tick the correct answer :

(i) The nucleus of an atom consists of:
(a) Electrons and protons.
(b) Protons and neutrons.
(c) Electrons and neutrons.
(d) None of these.

(ii) Which particle is the lightest in the following :
(a) Electron.
(b) Proton.
(c) Neutron.
(d) α-particles.

(iii) Which particle is heavier than others.
(a) Electron.
(b) Proton.
(c) Neutron.
(d) α-particles.

(iv) The mass of electron is:

(a) $9.11x10^{-26}$g. (b) $9.11x10^{-27}$g.

(a) $9.11x10^{-28}$g. (d) $9.11x10^{-30}$g.

(v) The mass of proton is:

(a) $1.67x10^{-22}$g. (b) $1.67x10^{-23}$g.

(c) $1.67x10^{-24}$g. (d) $1.67x10^{-25}$g.

(vi) Charge on an electron is:

(a) $1.6x10^{-16}$C. (b) $1.602\ x10^{-17}$C.

(c) $1.67x10^{-18}$C. (d) $1.602\ x10^{-19}$C.

3. Write answer of the following questions:

(i) Outline the main points of Daltons atomic theory.

(ii) What evidence is there that electrons are negatively charged particles?

(iii) Discuss Rutherford's gold metal foil experiment? What did it tell about the structure of the atom.

(iv) Explain the main features of Bohr's theory.

(v) What are protons and how were these produced?

(vi) What are distinguishing characters of electrons, protons and neutrons?

(vii) Define the following terms.

(a) $\propto$-particles (b) β-particles (c) γ -rays.

(viii) What is the proof that all atoms contain electrons?

(ix) Describe the three types of rays emitted by radio-active substances.

(x) Why it is believed that the atom has mostly empty spaces.

(xi) Define the terms:-

(a) Atomic number (b) Mass number (c) Isotopes.

(xii) In what way do isotopes of a given element differ form each other?

(xiii) A given isotope of nitrogen (N) contains 7 electrons, 7 protons and 8 neutrons.

(a) What is its mass number?

(b) What is its atomic number?

(xiv) C-14 and N-14 both have same mass number yet they are different elements. Explain.

(xv) What are the names of three sub-atomic particles? What are their masses in atomic mass units (a.m.u) and what is the unit charge on each?

(xvi) Give the names and symbols for the following elements.

(a) An element with atomic No. 6.

(b) An element with 18 protons in the nucleus.

(c) An element with 17 electrons.

(xvii) How many electrons and protons are there in each atom of the following?

(a) Carbon (b) Aluminium (c) Argon

(d) Fluorine (e) Potassium (f) Sulphur

(xviii) How many protons, neutrons and electrons are present in the following atoms?

(a) $^{15}_{7}N$ (b) $^{60}_{27}Co$ (c) $^{131}_{53}I$ (d) $^{207}_{82}Pb$

CHAPTER-4

PERIODICITY OF ELEMENTS

You will learn in this chapter about:

* The search for a classification.
* Dobereiner's classification.
* Newland's classification.
* Lother Meyer's classification.
* Mendeleev's classification.
* Salient Features of Mendeleev's Table
* Modern periodic table.
* Modern periodic Law.
* Periods and groups.
* Metals, non-metals and metalloids in the periodic table.
* Some periodic properties of atoms, like atomic radii, ionization energy, electron affinity and electronegativity.

4.1 THE SEARCH FOR A CLASSIFICATION

In the early era of science, only few elements were known, but new elements were being discovered. Hence, in order to facilitate their study, need arose for a frame work in which the elements may be arranged or classified. Previously the classification was based on atomic mass (atomic weight) of elements because it was thought that the properties of an element depended upon its atomic mass. But recently a complete classification has been made, which is based on atomic number of the elements, instead of their atomic masses.

We trace the history of development of classification of elements, as follows:

4.1.1 Dobereiner's Classification:

In 1829 Johann Dobereiner, noticed that of the three elements with very similar chemical behaviour i.e. Calcium (Ca), strontium (Sr) and Barium (Ba), the atomic mass of the middle element i.e. Sr is almost the arithmetic mean of the other two. This led him to call this group of three elements, a 'triad' and proposed the **Law or Rule of Triad,** which states that:

"Central atom of each set of triad had an atomic mass almost equal to the arithmetical mean of the atomic masses of other two elements".

Table 4.1 Triads

Elements		Atomic Mass	Mean Atomic Mass
Triad	Lithium Sodium Potassium	7 23 39	At. mass of Na $= \frac{7+39}{2} = 23$
Triad	Chlorine Bromine Iodine	35.5 81 126.5	At. mass of Br $= \frac{35.5+126.5}{2} = 81$

This law or rule cannot be extended to the classification of all the elements, because it is true only in the cases of very few elements only.

4.1.2 Newland's Classification:

In 1863 John Newland, a London industrial chemist proposed **NEWLAND'S LAW OF OCTAVE**, which states that:

"If elements are arranged in the order of increasing atomic masses, the eighth element starting from a given one, has similar properties as first one i.e. its properties are a kind of repetition of the first, like the eighth note in an octave of music."

Table 4.2 Octave Law

Element	Li	Be	B	C	N	O	F
Atomic Mass	7	9	11	12	14	16	19
Element	Na	Mg	Al	Si	P	S	Cl
Atomic Mass	23	24	27	28	31	32	35.5
Element	K	Ca					
Atomic Mass	39	40					

For example, Na is eighth element from Li and has similar properties, Mg is eighth element from Be and has similar properties, etc.

This arrangement of elements for the first time brought to light the existence of **PERIODICITY** i.e. recurrence of chemical and physical properties at regular interval and provided a great idea towards the development of modern periodic table. This law failed because it held good for the first sixteen elements but did not work after seventeenth element. Moreover hydrogen was not included in this sequence.

4.1.3 Lother Meyer's Classification:

Julius Lother Meyer, a German scientist, in December 1869 published a periodic table, in which the then known 56 elements were arranged on the basis of their atomic masses in nine vertical columns or groups from I to IX. But he laid down emphasis on the physical properties of elements.

Lother Meyer calculated the atomic volumes of elements. The atomic volume of an element is the volume which would be occupied by 1 gram atomic weight (1 mole) of atoms of element if it were a solid.

$$\text{Atomic volume} = \frac{\text{Gram atomic weight}}{\text{Density}}$$

In order to emphasize the concept of periodicity, he plotted a graph between atomic volumes of elements against their increasing atomic masses.

The curve obtained is shown in Fig 4.1. It consists of sharp peaks and broad minima.

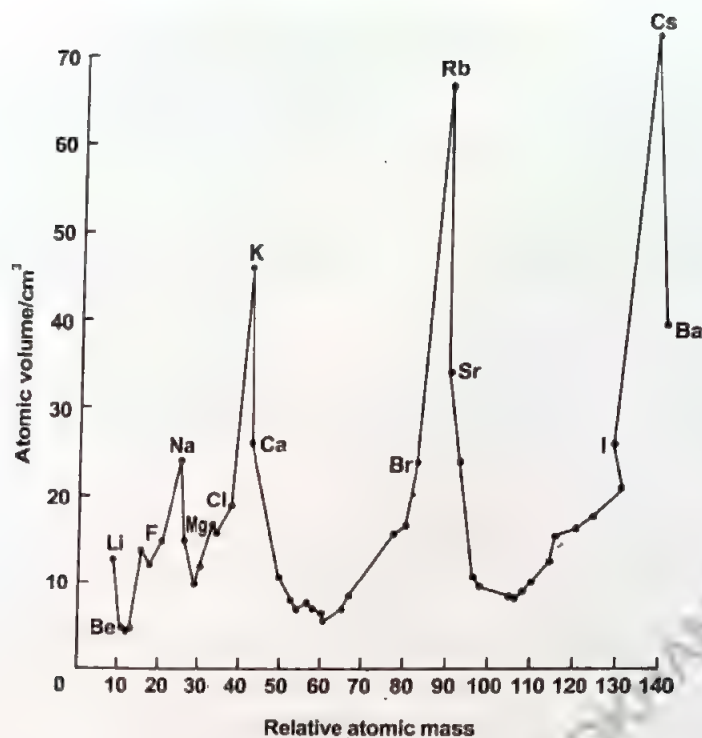



Fig 4.1 Lothar Meyer Atomic Volume Curve

He observed that the elements with similar properties occupy similar positions on the curve. For example, the highly reactive alkali metals (Li, Na, K, Rb, Cs) occupy the peaks there by showing that these elements have largest atomic volumes. The regular spacing of the highest points confirms the idea of periodicity, suggested by Newland.

4.1.4 Mendeleev's Classification:

In March 1869 Dimitri Mendeleev a Russian chemist arranged the elements in order of increasing atomic mass, placing the elements with similar chemical properties vertically beneath each other. By doing so, he observed that the properties of the elements with slight modification repeated themselves at intervals. So he put forward his Periodic Law which states that.

"The physical and chemical properties of elements are a periodic function of their atomic weights."

The periodic table published by Mendeleev consisted of eight vertical columns called groups (i.e. Group I to VIII) and horizontal rows called series. Now-a-days these series are called periods. The table is shown below:

Table 4.3 Mendeleev's Periodic Table of 1872

ROW	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
1	H = 1							
2	Li = 7	Be = 94	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27.3	Si = 28	P = 31	S = 32	Cl = 35.5	
4	K = 39	Ca = 40	___ = 44	Ti = 48	B = 51	Cr = 52	Mn = 55	Fe = 56 , Co = 59 Ni = 59 , Cu = 63
5	(Cu = 63)	Zn = 65	___ = 68	___ = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nh = 94	Mo = 96	___ = 100	Ru = 104 , Rh = 104 Pb = 106 , Ag = 108
7	(Ag=108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	I = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	...	...	...	
9	...	...	...	...	...	...	...	
10	...	...	?Er = 178	?La = 180	Ta = 182	W = 184	...	Os = 195 , Ir = 197 Pt = 198 , Au = 199
11	(Au = 199)	Hg = 200	T1= 204	Pb= 207	Bi = 208	...	...	
12	...	...	...	Th= 231	...	U = 240	...	

Spaces were left for the unknown elements with atomic masses 44, 68, 72 and 100.

Salient Features of Table:

1. It has eight vertical columns called groups and twelve horizontal rows called periods.
2. Elements in each vertical column have similar properties.
3. Vacant spaces were left for the elements not discovered until then. He proposed their names as eka-boron, eka-aluminium and eka-silicon.
4. The group number indicate the highest valence that can be attained by elements of that group.

Advantages of Mendeleev's Periodic Table:

1. It helped in systematic study of elements. For example the study of sodium helps means to a large extent in pridicting the properties of other alkali metals as potassium, rubidium, cesium. It forcefully proved the concept of periodicity.
2. Prediction of new elements was made possible.
   The physical and chemical properties of eka-boron, eka- aluminium and eka-silicon were predicted by Mendeleev. This helped in their discovery. These have been named as scandium (Sc), Gallium (Ga) and Germanium (Ge). Their properties are remarkably the same as were predicted by Mendeleev.
3. Mendeleev's periodic table helped in correcting many doubtful atomic masses.

INTERESTING TO NOTE

Discovery of Gallium and Confirmation of Mendeleev's Prediction

Gallium was unknown when Mendeleev developed his periodic table Mendeleev left many vacant spaces in his periodic table, stating that such vacant spaces would be filled by elements not yet discovered. He did not only left vacant spaces but he also predicted the properties of the elements which when found, would occupy these spaces, like those, eka (below) aluminium and silicon. In 1875, when the French Chemist Paul Emile Lecoq, discovered the new element Gallium (from Latin Gallia for France) and gave it density as 4.7g/cm^3. Mendeleev pointed out that the density should be 5.94g/cm^3, that was proved true as predicted. The scientific community was astounded that Mendeleev (Theorist) knew the properties of new element better than the chemist who has discovered it. The new element Gallium which was called by Mendeleev Eka-Aluminium, so named because it occupied the first vacant space in his periodic table under aluminium.

Mendeleev's periodic table proved to be the key for unlocking the mysteries of atomic structure and chemical bonding. In the nineteenth century, element 101, was named Mendeleevium (Md) in his honour.

Defects In Mendeleev's Periodic Table:

1. There are three pairs of elements i.e. elements of higher atomic masses placed before elements of lower atomic masses i.e.
   (a) Argon (40) placed before potassium (39)
   (b) Cobalt (59.9) placed before nickle (58.6)
   (c) Tellurium (127.6) placed before iodine (126.9)
2. No place for isotopes of elements.
3. Dissimilar elements placed in same group i.e. Alkali metals (Li, Na, K, Rb, Cs, Fr) were placed with coinage metals (Ag, Cu, Au)
4. Similar elements placed in different groups for example Barium (Ba) and lead (Pb) resemble in many properties but they are placed in separate groups.
5. It failed to give the idea of atomic structure.

4.1.5 Modern Periodic Table:

The arrangement of elements on the basis of their atomic masses left many anomalies in the position of different elements in the periodic table. Moreover the existence of isotopes showed that the atomic mass of an element is not the fundamental property of an element.

The modern periodic table is the result of discovery of atomic number by Moseley in 1914.

Based on the concept of atomic number Bohr, Werner and Bury proposed the **Modern Periodic Law** which states that,

"The physical and chemical properties of all elements are periodic functions of their atomic numbers."

In modern periodic table, also known as Bohr's Long Form Of Periodic Table, elements are arranged in order of their increasing atomic number. The elements having similar properties are repeated at regular intervals. As atomic number is related to the number of protons in an atom, so the real basis of periodicity of properties is due to recurrence of similar valency shell configuration of the next element in the same group. The modern periodic table is shown in table 4.4

Table 4.4 Modern Periodic Table of the Elements

Key	
1	Atomic number
H	Symbol
1.008	Atomic weight

Main groups 1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.00794																	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.81	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.998403	10 Ne 20.1797
11 Na 22.98977	12 Mg 24.305											13 Al 26.98154	14 Si 28.0855	15 P 30.97376	16 S 32.066	17 Cl 35.453	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.41	49 In 114.82	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.9045	54 Xe 131.29
55 Cs 132.9054	56 Ba 137.33	57 *La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.9665	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.9804	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.0254	89 †Ac 227.0278	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 Uun (269)	111 Uuu (272)	112 Uub (277)	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

Transition metals: groups 3–12. Main groups: groups 1–2 and 13–18.

Series														
*Lanthanide series	58 Ce 140.12	59 Pr 140.9077	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.9254	66 Dy 162.50	67 Ho 164.9304	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04	71 Lu 174.967
†Actinide series	90 Th 232.0381	91 Pa 231.0359	92 U 238.0289	93 Np 237.048	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

The modern periodic table contains seven horizontal rows called periods and sixteen vertical columns called groups.

4.1.6 Periods:

The elements within a period have dissimilar properties from left to right across any period, the physical and chemical properties of elements change from metallic to non metallic along a period.

All periods except the first, start with an alkali metal with one electron in their valence shell and end up with zero group element with valence shell having 8 electrons, except He which has only 2 electrons.

The First Period:

It contains only two elements i.e. H and He. This period signifies the completion of K-shell or first orbit. It is the shortest period with two elements.

The Second and Third Period (Short Periods):

Each of these periods contains 8 elements. They signify the filling up of L-shell and M-shell respectively.

The second period starts with Li and ends up with Ne; whereas the third period starts with Na and ends at Ar.

The Fourth and Fifth Period (Long Periods):

Each of these periods contains 18 elements. In these periods the electrons fill M and N shells. Fourth period starts from K and ends at Kr. Fifth period starts from Rb and ends at Xe.

The Sixth Period (Longest Period):

It contains 32 elements. It starts from Cs and ends with Rn. Besides, fourteen elements called **Lanthanides,** are placed at the bottom of periodic table.

The Seventh Period (Incomplete Period):

It starts with Francium (Fr). This period is incomplete as to date about 109 elements have been discovered.

This period also includes a group of fourteen elements starting from **Actinium** (Ac). These elements are called **Actinides**. They are also placed at the bottom of the table.

4.1.7 Group:

The vertical columns are called groups. Basically there are eight groups (I to VIII) but each group is further sub-divided into "A" and "B", sub-groups.

The elements of sub-group 'A' are called **'Main'** or **Representative Elements,** as the properties of these elements are represented by valency electrons.

The elements of sub-group 'B' are called **Transition Elements,** because the properties of these elements show a gradual change or transition between the two sets of representative elements, on either side of them.

Elements of a group have similar valency shell configuration hence have similar properties.

The group number indicates the total number of electrons in valency shell of that element.

Group I A (The Alkali Metals) or (Lithium Family)

This group includes Lithium (Li), Sodium (Na), Potassium (K), Rubidium (Rb), Caesium (Cs) and Francium (Fr). Their valence shell contains one electron only, and on reaction they lose this electron and form univalent positive ions (M^{1+}). They are highly reactive metals with low melting points. Fr is radioactive.

Their atomic radii, atomic volumes, ionic radii increase from Li to Cs due to the addition of extra shell to each element and due to same reason, the melting and boiling points decrease downward. They are called **Alkali Metals** because they form water soluble base such as NaOH and KOH.

Group IIA (The Alkaline Earth Metals); (Beryllium Family)

It includes Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium(Sr), Barium (Ba) and Radium (Ra).

Their valence shell contains two electrons. On reaction they lose these two electrons and form divalent positive ions (M^{2+}). Ra is radioactive.

These elements are a bit harder, having higher melting and boiling points than the alkali metals, but they have smaller atomic, ionic radii and atomic volumes.

Down the group they do not show a regular trend in melting, boiling points and densities.

Group IIIA (The Boron Family):

It includes Boron (B); Aluminium (Al); Gallium (Ga); Indium (In) and Thallium (Tl). Their valency shell contains three electrons. They exhibit a valency of 3^+ and form M^{3+} ions. Except boron they are highly electropositive elements i.e. having metallic character which increases down the group; due to increase in atomic volume. Boron is **Metalloid**. A metalloid is an element which has some properties of metals and some properties of non-metals.

Group IVA (Carbon Family):

It includes Carbon (C); Silicon (Si); Germanium (Ge), Tin (Sn) and Lead (Pb).

Their valence shell contains four electrons. C, Si and Ge form covalent compounds whereas Sn and Pb exhibit a variable valence of 2 and 4.

Of these elements C is non-metal; Si and Ge are metalloids, Sn and Pb are metals.

Down the group atomic radii and volumes increase due to addition of a new shell and for the same reason metallic character increases down the group. C and Sn exist in different allotropic forms.

Group V (Nitrogen Family):

It includes Nitrogen (N), Phosphorus (P) Arsenic (As), Antimony (Sb) and Bismuth (Bi).

Of these elements N and P are non-metals, As and Sb are metalloids and Bi is a metal. Their valence shells contain five electrons. There is a large variation of properties as we go down the group.

Nitrogen exists as diatomic molecules (N_2) and forms a number of oxides as NO, N_2O, NO_2, N_2O_4 and N_2O_5. Due to small atomic size and large ionization potential, nitrogen has a tendency to accept three electrons to form nitride ion (N^{3-}). Phosphorus exists as P_4 molecule.

Except Nitrogen all exist in more than one allotropic forms.

Group VI (Oxygen Family):

It includes oxygen (O); Sulphur (S); Selenium (Se), Tellurium (Te) and Polonium (Po).

Of these oxygen and sulphur are non-metals, selenium, tellurium are metalloids and polonium is metal.

All of the elements exhibit the property of allotropy. For example

allotropic forms of oxygen are oxygen (O_2) and ozone (O_3).

Oxygen and sulphur form divalent negative ions O^{2-} and S^{2-}. Their valence shell contains six electrons.

Group VIIA (The Halogens):

It includes Fluorine (F); Chlorine (Cl); Bromine (Br); Iodine (I) and Astatine (At).

Except Astatine (which is a metalloid) all others are non-metals and exist as diatomic molecules. At room temperature F_2 and Cl_2 are gases; bromine is a liquid and iodine is a solid. Their valence shell contains seven electrons. They have high ionization energies and large negative electron affinities hence they easily accept an electron to form halide ions $(x)^{1-}$ i.e. (F^{1-}, Cl^{1-}, Br^{1-}, I^{1-})

Group VIIIA (Inert or Noble Gases) :

It includes Helium (He); Neon (Ne); Argon (Ar); Krypton (Kr); Xenon (Xe) and Radon (Rn).

Their valence shell contains eight electrons, except helium which has two electrons. With the exception of krypton and xenon (which have large atomic volumes so slightly reactive under drastic conditions) the rest of these elements are totally inert chemically. The reason is that these have completely filled outer shells, a condition that represents greater stability.

INTERESTING TO NOTE

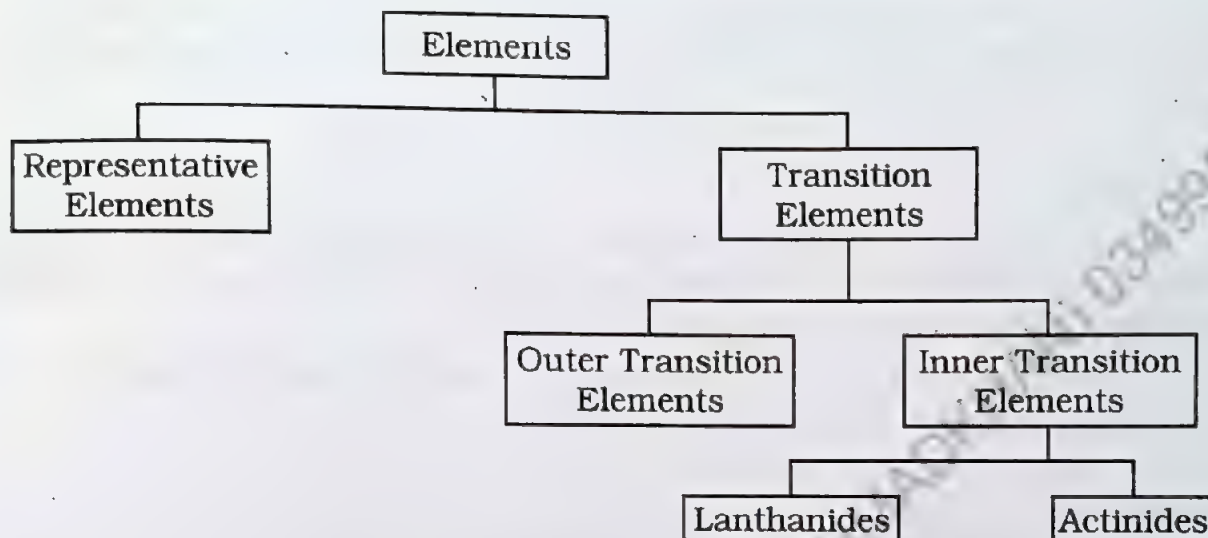
Discovery of Noble Gases

None of the noble gases was known when Mendeleev proposed his periodic table. In 1892. The English scientist Ramsay became interested in the discovery that nitrogen obtained from the air had a slightly higher density than that prepared by chemical reactions. After careful investigation, he concluded that higher density must be due to the presence of unknown gas. When he separated this gas from the air, he found that it was completely unreactive. He called it argon, the "idle or lazy" gas in Greek.In the same year Ramsay isolated helium (He) the lightest of all noble gases, from uranium ores. During 1898 Ramsay and Rayleigh isolated three additional noble gases from air, neon (Ne) krypton (Kr) and xenon (Xe).

Group IB To VIIIB (Transition Elements):

These are metals. In these elements, besides the valence shell penultimate shell is also incomplete. In chemical reactions they show more than one valencies. These elements in compounds having characteristic colours.

The total classification can be summarized in the following scheme:

- Elements
  - Representative Elements
  - Transition Elements
    - Outer Transition Elements
    - Inner Transition Elements
      - Lanthanides
      - Actinides

4.1.8 Non-metals, Metalloids and Transition Metals in Periodic Table.

Majority of the known elements are metals; only 17 elements are non-metals and eight elements are metalloids.

Table 4.5

Metals (IA–IIA); METALS (IIIB–IIB); Non Metals and Metalloids (IIIA–VIIIA)

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Te	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une									

METALS

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Metals:

They are electropositive elements i.e. they lose electrons to form cations. They form basic oxides.

All of them have lustre and are, malleable (i.e. can be spread out into sheets) and ductile (i.e. can be drawn into wire), are good conductors of heat and electricity.

In the periodic table, elements of group IA, IIA and all transition elements are metals. Some of the elements of group IIIA, IVA, VA and VIA are also metals.

Non-Metals:

They are electronegative elements i.e. they gain electrons to form anions. They form acidic oxides. They are bad conductors of heat and electricity. Most of them are gases.

In the periodic table, majority of elements of p- block i.e. group IIIA, IVA, VA, VIA, VIIA and VIIIA are non-metals.

Metalloids:

These are the elements which exhibit dual character. That is they show the properties of both metals as well as non-metals. For example, their oxides are amphoteric i.e. have basic as well as acidic nature.

Examples are:
Boron (B) of group IIIA
Silicon (Si) and Germanium (Ge) of group IVA.
Arsenic (As) and Antimony (Sb) of group VA.
Tellurium (Te) and Polonium (Po) of group VIA
Astatine (At) of group VIIA.

4.2 SOME PERIODIC PROPERTIES OF ATOMS

4.2.1 Atomic Radii:

Modern research shows that an atom does not have strictly defined boundaries. So it is impossible to determine the exact radius of an atom, but for all practical purposes the atomic radius may be defined as half the distance between two adjacent nuclei of two similar atoms in touch with each other.

It is measured in Angstrom unit (A^0 or A.U)
$1A^0 = 10^{-8}$cm.

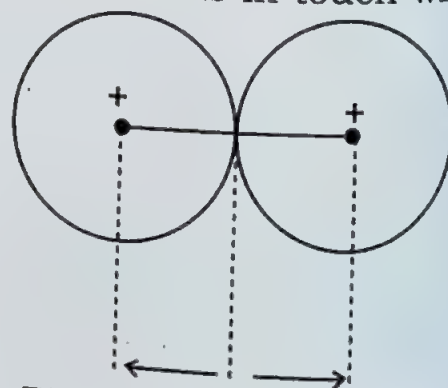



Fig. 4.2 Atomic Radius

The atomic radius depends upon the number of shells and nuclear charge in an atom.

In the periodic table the atomic radii increase down the group due to addition of new shell in each atom. But in a period the atomic radii decrease from left to right due to increase in number of protons i.e. increase in nuclear charge, which results in stronger pull on orbiting electrons by the nucleus. The variation of atomic radii in a group and a period is shown is following tables.

Table 4.6 Atomic radii Down a Group.

Elements of Group IIA	Atomic Radii in A^0
Be	1.12
Mg	1.36
Ca	1.97
Sr	2.15
Ba	2.22

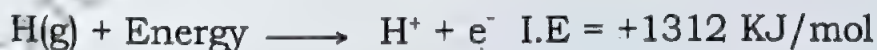
Table 4.7 Atomic Radii in a Period

Elements	Na	Mg	Al	Si	P	S	Cl
Atomic Radii in A^0	1.51	1.36	1.25	0.77	0.70	0.66	0.64

4.2.2 Ionization Energy:

Ionization energy is one of the few fundamental properties which can be measured directly. It is defined as the minimum energy required to remove an electron from a gaseous atom in its ground state. It is measured in K.Jole/mole or electron volt (ev)per atom.

Ionization energy depends upon atomic size and nuclear charge. The higher the ionization energy the more difficult is to remove an electron. The ionization energy of hydrogen is 1312 K.J/mol. i.e.

$$H(g) + \text{Energy} \longrightarrow H^+ + e^- \quad I.E = +1312\ KJ/mol$$

Down a group in the periodic table, the ionization energy decreases because the addition of a new shell decreases the hold of nucleus on valence electron.

Ionization energy increases from left to right in a period because

the addition of protons in the nucleus, increases the nuclear charge there by increasing the force of attraction on electrons.

The amount of energy required to remove first electron is called **first ionization energy.** For subsequent electrons it is called second, third, fourth ionization energy, etc.

The periodicity of ionization energy can be observed in the plot between first ionization potential and atomic number as shown in Figure 4.3.

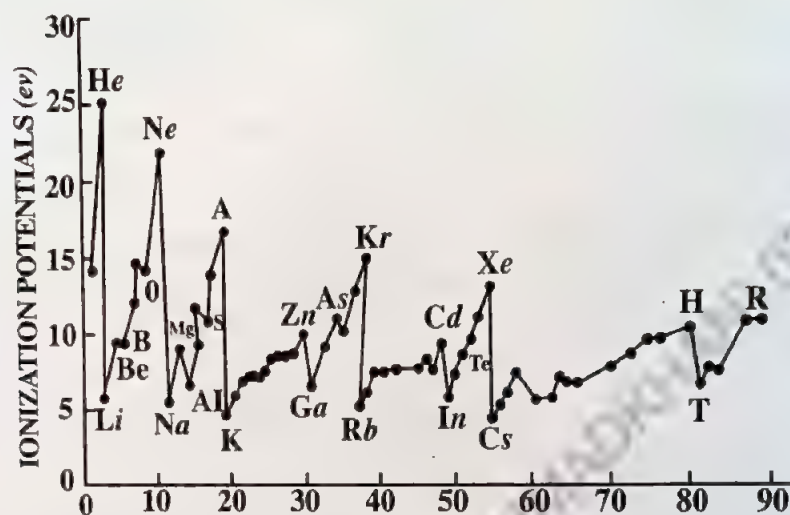
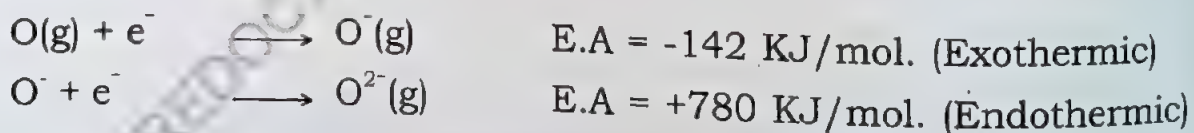



Fig. 4.3 Atomic number

4.2.3 Electron Affinity:

Electron affinity is defined as the energy change that occurs when an electron is gained by an atom in the gaseous state. It is measured in KJ/mol or in e.v per atom. Electron affinity for the addition of first electron is negative i.e. Energy is released but for further addition of electrons it is positive, because energy has to be added to overcome repulsion between negative ion and electron, as shown below:

$O(g) + e^- \longrightarrow O^-(g)$ E.A = -142 KJ/mol. (Exothermic)

$O^- + e^- \longrightarrow O^{2-}(g)$ E.A = +780 KJ/mol. (Endothermic)

Electron affinity depends upon the atomic size and nuclear charge.

Down a group in the periodic table, electron affinity decreases because the addition of a new shell to each atom decreases its force of attraction.

Table 4.8

Element	Electron Affinity in KJ/mol
F	-333
Cl	-348
Br	-324
I	-295

Flourine (F) has abnormally low electron affinity because due to its very small atomic size it does not accept electron easily.

In a period, the electron affinity increases from left to right because successive atoms have higher nuclear charge and attract the incoming electron more towards itself. See table 4.9

Table 4.9

Element	Li	Be	B	C	N	O	F	Ne
E.A in KJ/mol	-58	0	-23	-123	0.2	-142	-333	O

4.2.4 Electronegativity (E.N):

Electronegativity is defined as the relative tendency of an atom in a molecule to attract shared pair of electrons to itself. It is denoted by a number and has no unit.

Linus Pauling calculated the electronegativities of different elements taking fluorine as standard with its electronegativity = 4

Down a group electronegativity decreases as due to addition of new shell, the power of a nucleus to attract electron decreases. In a period from left to right it increases due to increase in nuclear charge.

Generally speaking, elements with high ionization energy and large electron affinity have high electronegativity. Variation of electronegativity with atomic number and the periodicity in it is shown in the following figure.

Electronegativity

4
3
2
1

F
Cl
Br
Rb
Ru
I
Mn
Zn
Li
Na
K
Rb

0
10
20
30
40
50

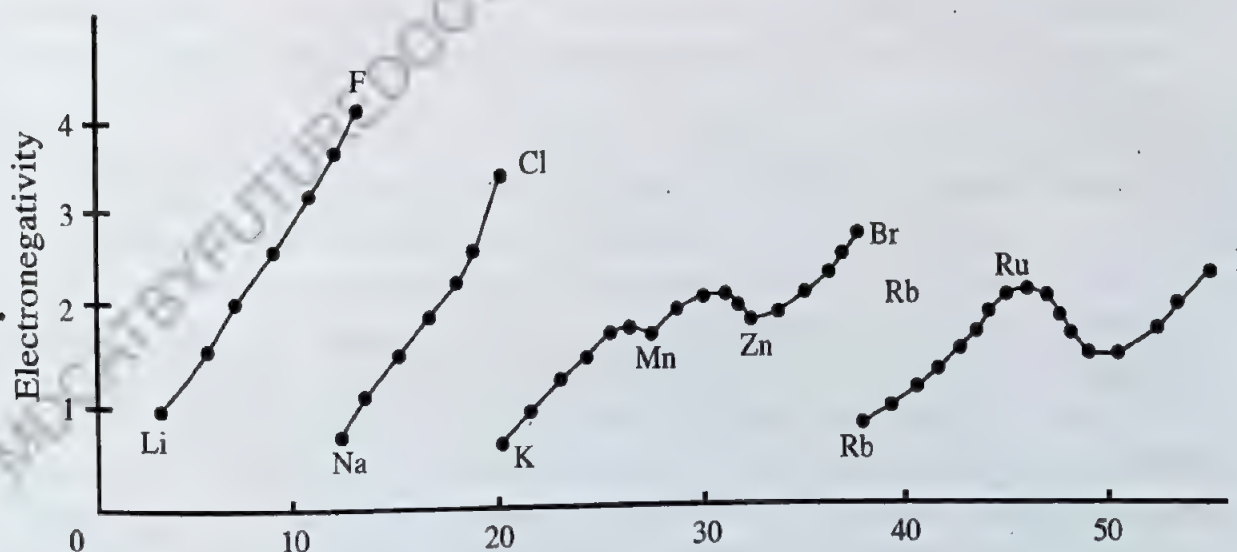



Fig 4.4 Atomic Number

The difference in electronegativities of two combining atom decides the nature of bond that is formed between them and affects the properties of molecules.

SUMMARY

1. With the discovery of new elements, scientists tried to arrange them in a scientific and organised manner, on the basis of their atomic masses. Dobereiner in 1829 was the first to classify the similar elements in the groups of three elements, known as **"triads"**. Newland in 1863 then classified the elements in the increasing atomic masses and put forward the **law of octaves.** In 1869 German chemist Lothar Meyer improved the Newland's classification and plotted a graph between the atomic volumes and atomic masses of the elements. He included only 56 elements and discovered that the elements with similar properties occupied the similar positions on the curve.
2. In 1869 Russian scientist Mendeleev arranged the elements in the increasing order of atomic masses and put forward a periodic law, known as Mendeleev's periodic law. In 1873, he published his periodic table which consisted of eight vertical columns, known as groups and twelve (12) horizontal rows known as periods. Many places were left vacant for undiscovered elements in his periodic table and predicted their properties. He also discovered correct atomic masses of many elements.
3. In (1913-1914), after the discovery of atomic numbers, the elements were then arranged in their increasing atomic numbers and modern periodic law was introduced.
   According to modern periodic law. "The physical and chemical properties of the elements are the periodic functions of their atomic numbers.
4. According to modern periodic classifications, a long-form of periodic table has been introduced. It contains eight (8) groups and seven (7) periods. The groups are further subdivided into A and B sub-groups.
5. In the periodic table, the properties of the elements gradually change in groups and periods. Down the groups generally the physical properties of the element slightly change with the increasing atomic numbers, while chemical properties remain same to some extent. Along the periods the physical and chemical properties both change to larger extent with the increasing atomic numbers. The chemical properties depend upon the valence electrons present. The elements with similar electronic configurations in their valence shells have similar chemical properties.

6. In the periodic table, elements on the left are metals, while on the right side are non-metals and in the middle, the elements are generally metalloids showing both the characteristics of metals and non-metals.
7. The minimum amount of energy required to remove the most losely bound electron from the valence electrons in a gaseous atom to form positive ion. i.e. cation is called first ionization energy or ionization potential and the minimum amount of energy released from a gaseous atom when an electron is added in the valence shell to form negative ion i.e. anion is called Electron affinity. Down the groups the values of ionization potential and electron affinity both decrease with the increasing atomic numbers due to increase in atomic sizes. While along the periods the values of ionization potential and electron affinity both increase with the increasing atomic numbers.
8. The relative power of an atom to attract the shared pair of electrons towards itself is called electronegativity. The most electronegative atom is that of fluorine (F) with electronegativity = 4.0.
   The repetition or reccurrence of properties after regular intervals is called periodicity.

EXERCISE

1. Fill in the blanks:

(i) The rule of triad was introduced by ....

(ii) The repetition of properties after regular intervals is called ......

(iii) The longest period is .... period and contains total ..... elements.

(iv) The elements that contain both metallic and non-metallic characteristics are called ....

(v) The long form of periodic table contains ..... groups and ...... periods.

(vi) According to Mendeleev the properties of the elements are the periodic functions of their ....

2. Tick True or False in the following statements:

(i) Mendeleev put forward his periodic law in 1856.

(ii) The first period contains two elements, hydrogen and hellium.

(iii) The longest period in the periodic table is 7th period.

(iv) Lanthanides and Actinides are d-block elements.

(v) Down the group the electronegativity increases with increasing atomic number.

(vi) The law of octaves was introduced by John Newland.

(vii) Li^7, Na^{23} and K^{39} form a triad.

3. Pick up the correct answer: (Multiple Choice Questions).

(i) Mendeleev's periodic table contained .... periods
(7, 8, 12, 10)

(ii) The incomplete period in the periodic table is ....
(7, 6, 3, 1)

(iii) The most reactive metal is ....
(Na, Cu, Fe, Ca)

(iv) The only liquid metal is ......
(Molybdenum, Gold, Mercury, Bromine)

(v) Lothar Meyer's curve included about ...... elements.
(Thirty, forty, fifty six, sixty two)

(vi) To which family does Ga belong ?
(Boron, Carbon, Nitrogen, Fluorine)

(vii) The elements of VII A group are known as,
(Halogens, Lanthanides, Actinides)

4. Write answer to the following questions:

(i) Define the followings:
(a) Doberneir's rule of triad (b) Periodicity
(c) Modern periodic law (d) Electronegativity

(ii) If an element contains two shells only and its outershell contains five electrons then to which group the element belongs in the periodic table. Name the element. Predict its period.

(iii) State Mendeleev's periodic law. Describe Mendeleev's periodic table. Write down the advantages and disadvantages of Mendeleev's periodic table.

(iv) Explain Newland's law of Octave. How this law provided the larger scope for the classification of the elements?

(v) Which pair of elements is chemically similar?
(a) K, Cr (b) Cu, Ca
(c) F, Cl (d) N,O

(vi) What do you understand by long form of periodic table? Explain some of its applications?

(vii) Discuss some of the physical properties of the elements which exhibit periodicity.

(viii) The elements having eight valence electrons are known as,
(a) Noble or inert gases (b) Halogens (c) Nitrogen family
(d) Transition elements.

(ix) How does the modern periodic law differ from Mendeleev's periodic law? Explain clearly groups and periods in the modern periodic table.

(x) What do you understand by the periodic classification of elements? What are the merits and demerits of the classification of the elements in the periodic tabular form?

(xi) What do you understand by representative and transition elements?

(xii) Discuss the following physical properties of the elements.
(a) Ionization energy (b) Electron affinity
(c) Electronegativity (d) Atomic Radii

(xiii) What are lanthanides and actinides? Are they d or f-block elements?

CHAPTER-5

CHEMICAL BONDING

You will learn in this chapter about :

- Why do atoms form chemical bonds.
- Ionic bond.
- Characteristics of ionic compounds.
- Covalent bond.
- Characteristic of covalent compounds.
- Electronegativity values.
- Ionic character of covalent bond.
- Co-ordinate covalents bond or Dative covalent bond.
- Difference between covalent and Co-ordinate covalent bonds.
- Metallic bonding.
- Inter-molecular forces.
- Hydrogen bonding.

5.1 INTRODUCTION

As we examine the world around us, we find that it is generally made up of compounds and mixtures.Rocks, coal, soil, air, trees, animals, all are formed by combination of atoms. Substances composed of single atoms are very rare. Examples are, argon (Ar) in the atmosphere and helium (He), mixed with natural gas. Clearly, there must be some force that holds atoms together in a molecule or crystals, otherwise the atoms would simply fly apart, and no compound could exist. *The force which holds atoms together in a molecule or a crystal, is called a chemical bond.*

5.2 FORMATION OF CHEMICAL BONDS.

Before the discovery of electronic structure of atoms, the nature of the forces holding the atoms together in a molecule or crystal was a mystery. Now it is believed that these forces are electrical in nature and that the chemical reactions, that occur between atoms involve change in their electronic structures. The electrons in the outer most shell of an atom are called the valence electrons. In the formation of chemical compounds from the elements, the valence electrons are either transferred from the outer shell of one atom to the outer shell of another atom or shared between them.This produces a chemical bond. When an atom of one element chemically combines with the atom of another element, both atoms usually attain a stable outer shell, consisting of eight electrons (octet). Only hydrogen H_2 and helium (He) atom have the stable outer shell of two electrons (Duplet). This is in accordance with the general rule that all processes tend to move towards the state of maximum stability. Generally, a stable molecule occurs, when the total energy of the combined atoms, is less than the total energy of the individual atoms.

5.3 TYPES OF CHEMICAL BONDS.

The first explanations of the nature of chemical bonds were advanced by W.Kossel (a German scientist) and G. N. Lewis (an American chemist) in 1916. They proposed two major types of chemicals bonds.

1. The ionic or electrovalent bond (by the transfer of one or more electrons from one atom to another, to form ions).
2. The covalent bond (a bond that results when atoms share electrons). Other types of bond include metallic bonds and hydrogen bonding.

> ***REMEMBER***
>
> *Chemical bonding also plays a role in determining the state of matter. At room temperature, water is liquid, carbon dioxide (CO_2) is a gas and table salt, sodium chloride (NaCl) is solid, because of difference in chemical bonding.*

5.4 IONIC BOND OR ELECTROVALENT BOND

In this type of combination, there is a complete transfer of one or more electrons from one atom to another. The atom that transfers electrons gets positive charge and the atom that gains electrons gets negative charge . The strong electrostatic force acting between positive (+ve) and negative (-ve) ions, holds them together. *The attraction that binds oppositely charged ions together is termed electrovalent bond or ionic bond.*

For illustration, let us consider the combination of sodium (Na) and chlorine (Cl) atoms to form common salt, sodium chloride (NaCl). In this combination, an atom of Sodium (Na) transfers one outer most shell electron and becomes positive sodium ion (Na^+) and an atom of chlorine gains that one electron to complete its octet and becomes chloride negative ion (Cl^-).

$$\underset{2,8,1}{Na} \longrightarrow \underset{2,8}{Na^+} + e^-$$

$$\underset{2,8,7}{Cl} + e^- \longrightarrow \underset{2,8,8}{Cl^-}$$

$$Na^+ + Cl^{1-} \longrightarrow Na^+ \;\; Cl^- \text{ or } Na^{1+}Cl^{1-}$$

Na   Cl $Na^+ Cl^-$

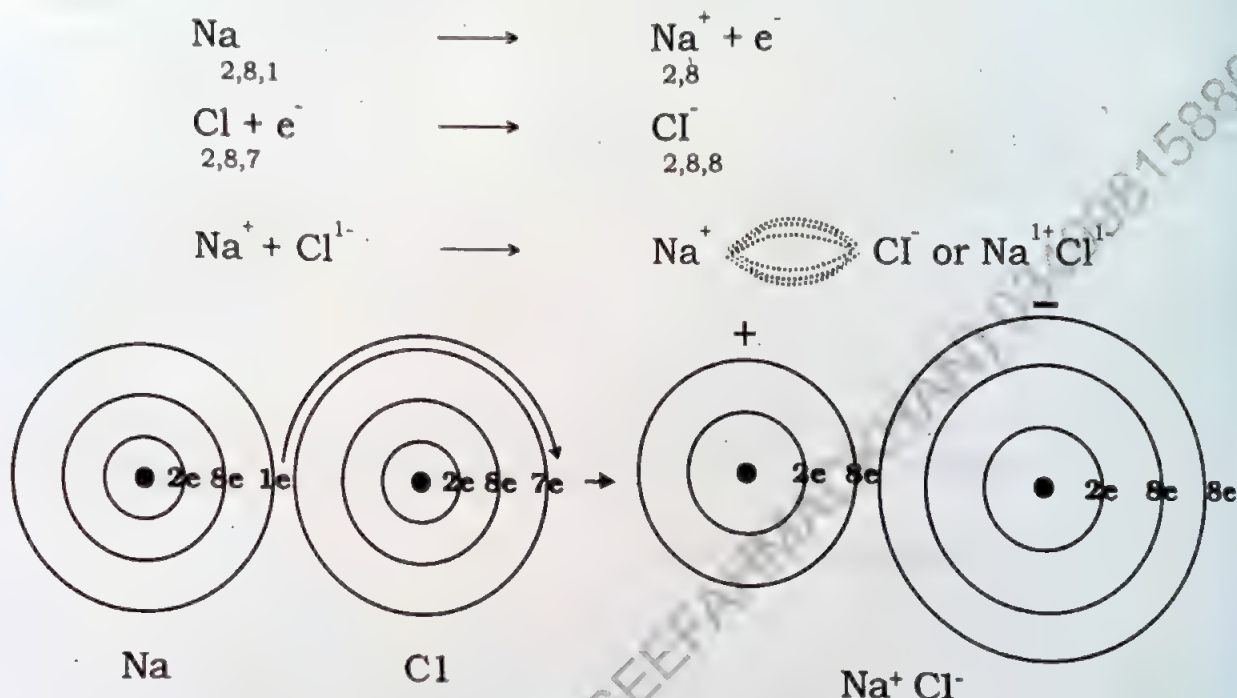



Fig 5.1 Formation of sodium chloride

The attraction that binds (Na^+) and (Cl^-) ions together is called electrovalent bond and the compound (NaCl) is called electrovalent compound or ionic compound.

IMPORTANT TO NOTE

The electrovalent bond is also known as ionic bond because electrovalent compounds when dissolved in water or melted, produce ions and therefore, conducts electricity.

Another example of electrovalent bond, is the formation of magnesium oxide (MgO) from magnesium and oxygen. The magnesium (Mg) atom has two electrons in its valence shell. If these two electrons are lost, the (Mg) will become a di-positive ion (Mg^{+2}), and the oxygen (O) atom gains these two electrons to complete its octet, and becomes di-negative ion (O^{-2}). These two oppositely charged ions form electrovalent bond and the compound (Mg O) is called electrovalent compound.

$$\underset{2,8,2}{{}_{12}Mg} \quad \underset{2,6}{{}_{8}O} \rightarrow \underset{2,8}{Mg^{+2}} \quad \underset{2,8}{O^{-2}} \longrightarrow Mg^{2+}O^{2-}(s)$$

2e 8e 2e   2e 6e →   2e 8e   2e 8e

Mg   O Mg^{++} ion O^{--} ion

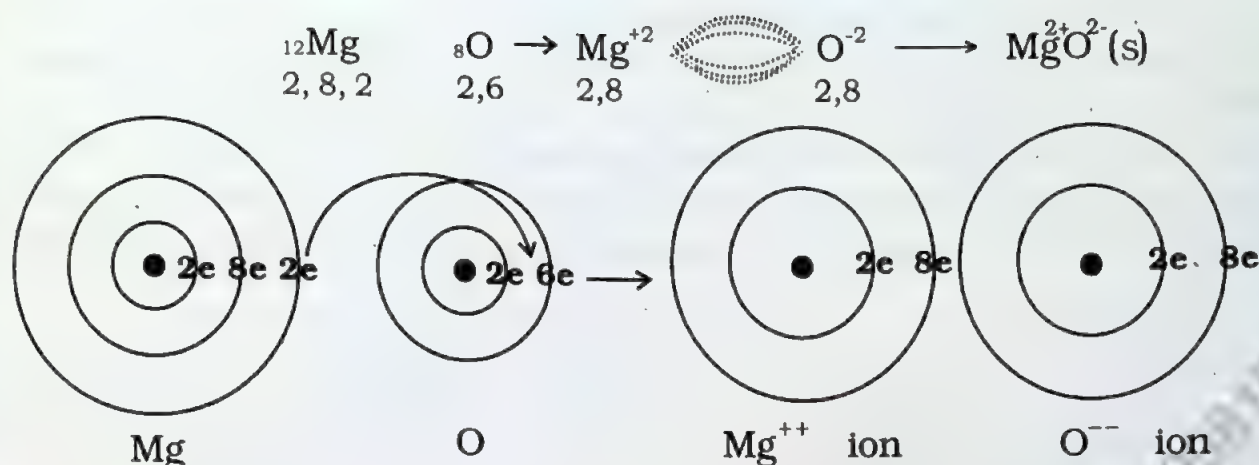



Fig. 5.2 Formation of magnesium oxide

5.4.1 Characteristics of Ionic Compounds:

1. In ionic bond, it is impossible to say that any two ions are bonded to each other to produce molecule, but in the crystals of ionic compounds, the oppositively charged ions are mutually surrounded by each other in orderly arrangement. Thus ionic compounds are solids at room temperature.
2. Ionic compounds, have high melting and boiling points because of the strong electrovalent bonds existing between the ions.
3. Ionic solids, do not conduct electricity as the ions are not free to move. Once an ionic compound is melted (fused) the ions are free to move and conducts electricity. Similarly, solutions of ionic compounds conduct electricity.
4. Ionic compounds are usually soluble in polar solvents, i.e. solvents of high di-electric constant such as water. But ionic compounds are insoluble in non polar (organic) solvents. These solvents have low di-electric constant such as benzene, carbon tetrachloride, etc. They are mostly inorganic compounds.

5.5 COVALENT BOND

When two or more atoms of the same element or atoms of different elements having similar electronegativities react, the transfer of electrons does not occur. In these instances, the atoms achieve inert gas (noble gas) structure by sharing of electrons. Thus, the atoms complete their outer most shell by means of sharing of unpaired electrons, and a covalent bond is formed. In covalent bond each atom has to contribute equal number of unpaired electrons. *The shared pair of electrons which links the atoms in a molecule is known as covalent bond.*

In covalent bond the shared electron pair is commonly expressed by single short line (—) For example, the halogens (chlorine atoms) posses an electronic configuration in which there are seven electrons in their outer most shell, and lacking only one electron in order to attain the structure of an inert gas. This structure may be attaine by the halogens molecule (chlorine molecule) when both atoms share one electron for bond formation.

17Cl
2,8,7

$\rightarrow$ Cl – Cl

(Shared pair of electrons)
or
covalent bond

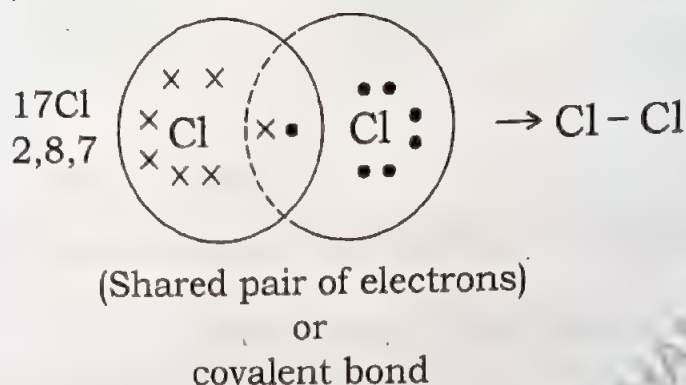



Fig. 5.3 Formation of molecular chlorine Cl_2

5.5.1 Single, Double and Triple Covalent Bonds:

(i) Single Covalent Bond:

In single covalent bond only one pair of electrons is shared by the bonded atoms, in which each atom has to share one electron. This type of bond is represented by single short line (—). For example, in the formation of (H_2) and hydrogen chloride (HCl) molecules, only one pair of electrons is shared.

i.e. H – H

i.e. H – Cl

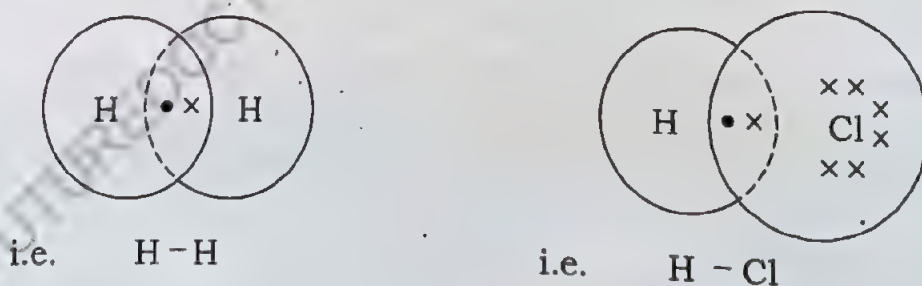



Fig. 5.4 Formations of single covalent bonds in H_2 and HCl molecules

(ii) Double Covalent Bond:

In double covalent bond only two pairs of electrons are shared by the bonded atoms, and each atom has to share two unpaired electrons. This type of bond is represented by two short lines (=) as shown in the molecules of oxygen (O_2) and carbon dioxide (CO_2).

i.e. $O = O$

i.e. $O = C = O$

Fig. 5.5 Formations of double covalent bonds in O_2 and CO_2 molecules

(iii) Triple Covalent Bond:

In triple covalent bond only three pairs of electrons are shared between the bonded atoms, and each atom has to share three unpaired electrons. This type of bond is denoted by three short lines ($\equiv$) as shown in the molecules of nitrogen (N_2) and ethyne (C_2H_2).

$N \equiv N$

$H - C \equiv C - H$

Fig. 5.6 Formations of triple covalent bonds in N_2 and C_2H_2 molecules

5.5.2 Characteristics of Covalent Compounds:

1. Compounds with covalent bonds are usually made up of discrete units (molecules) with a weak inter molecular forces.
2. In the solid state, there are weak vander wall forces between the molecules. Hence covalent compounds are often gases, liquids or soft solids with low melting points. In few cases, three dimensional covalent structures are formed rather than discrete units, hence diamond and silica (SiO_2) are covalent but are very hard and have high melting points. Usually covalent compounds have low melting and boiling points.
3. They are insulators because they do not conduct electricity.
4. Covalent compounds are usually insoluble in polar solvents like water, but soluble in organic solvents like benzene, ether, carbon tetra chloride etc.

5.5.3 Electronegativity: (E. N.)

If the covalent bond is formed between two like atoms, that molecule is called non-polar because the electron pair is shared equally between the two atoms, as in case of (H-H), (Cl-Cl), (O=O) and (N≡N) molecules. However, if the covalent bond is formed between the two dissimilar atoms as in hydrogen chloride (H-Cl) molecule, the attraction for electron pair, would not be equal, one atom will attract more than the other. Hence the electron pair will be displaced from the central position and reaches near to the chlorine atom.

This power of an atom to attract the shared pair of electrons towards, itself, is known as electronegativity.

Consequently, the chlorine atom being more electronegative tends to be partially negative and hydrogen atom would be partially positive.

$$H^{\delta+} — Cl^{\delta-}$$

Table 5.1 Values of Electronegativities of Different Elements.

Values of E.N of Different Elements

(Key: Atomic number / H Symbol / 1.008 Atomic weight)

IA	IIA	IIIB	IVB	VB	VIB	VIIB	IVB			IB	IIB	IIIA	IVA	VA	VIA	VIIA
Li 1.0	Be 1.5											[illegible]	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	[illegible]	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	[illegible]	[illegible]	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	[illegible]	[illegible]	I 2.5
Cs 0.7	Ba 0.9	La-Lu 1.1-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	[illegible]	[illegible]
Fr 0.7	Ra 0.9	Ac-No 1.1-1.7														

The table 5.1 shows a scale of electronegativities devised by Linus Pauling.

These values are based upon an arbitrary scale, in which fluorine is given an arbitrary standard value of electronegativity as 4.0. It is the most electronegative element. The electronegativity values of other elements are compared with that of fluorine. Note that the non metals have higher electronegativity values than the metals. Fluorine (F) has the highest electronegativity (4.0) and the metal cesium (Cs) has the lowest electronegativity (0.7) value.

5.5.4 Ionic Character in a Covalent Bond:

If covalent bond is formed between two like atoms e.g. in case (H-H), the hydrogen atoms are identical, hence the shared pair of electrons is not disturbed from the centre. This molecule is called ***non-polar*** because it is electrically neutral as well as symmetrical.

If covalent bond is formed between dissimilar atoms e.g. in case of H-Cl molecule, the shared pair of electrons move closer to one of the both atoms. In H-Cl molecule chlorine is more electronegative, hence the shared pair of electrons, is drawn nearer to chlorine than hydrogen atom. This results in partial positive charge on H atom and partial negative charge on Cl atom.

$$\overset{\delta^+}{H} — \overset{\delta^-}{Cl}$$

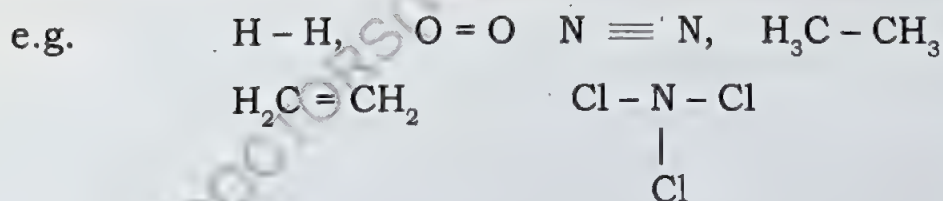
The covalent bond between H and Cl in H-Cl is partially ionic or polar covalent bond, because of positive and negative charges or poles.

Covalent bonds are partially ionic, if they exist between two dissimilar atoms and their ionic character depends upon the difference in electronegativities of bonded atom.

(i) Non-Polar Covalent Bond:

According to the scale of Linus Pauling, if the difference in the electronegativities of bonded atoms is zero, then the bond is pure covalent bond or non-polar bond. The molecule containing like atoms or the atoms of same electronegativities form a pure covalent bonds or non-polar bonds.

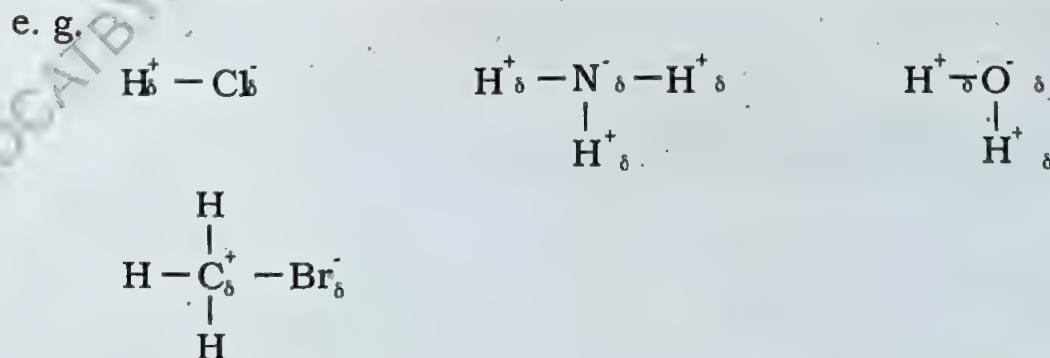
e.g. $H - H$, $O = O$, $N \equiv N$, $H_3C - CH_3$

$H_2C = CH_2$, $Cl - N(Cl) - Cl$

(ii) Polar Covalent Bond:

If the difference in the electronegativities of bonded atoms is upto 1.7 that bond is called polar covalent or partially ionic in character. The elements of different electronegativities always form polar covalent bond.

e. g.

$H^{\delta+} - Cl^{\delta-}$, $H^{\delta+} - N^{\delta-}(H^{\delta+}) - H^{\delta+}$, $H^{\delta+} - O^{\delta-}(H^{\delta+})$

$H - C^{\delta+}(H)(H) - Br^{\delta-}$

(iii) Electrovalent Bond:

If the difference in the electronegativities of bonded atoms is more than 1.7 then that bond is purely ionic or electrovalent. The bond between sodium and chlorine in the common salt (NaCl) is clearly ionic, because the difference in the electronegativities is 2.1 i.e. more than 1.7.

Na .9   Cl 3.0

2.1

$:\ddot{\underset{..}{Cl}}:\ddot{\underset{..}{Cl}}: \longrightarrow \overset{\delta+}{H}\ :\overset{\delta-}{\ddot{\underset{..}{Cl}}}: \longrightarrow Na^{+}\left[:\ddot{\underset{..}{Cl}}:\right]^{-}$

Non-polar covalent bond   Polar covalent bond   Ionic bond

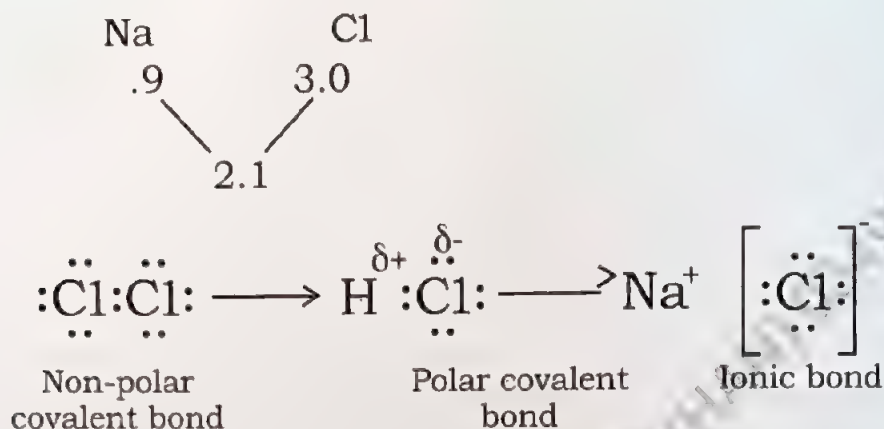



Fig: 5.7 Illustrates non-polar covalent bond (Cl_2),

Polar covalent bond (HCl) and Ionic bond (NaCl)

5.6 CO-ORDINATE COVALENT BOND OR DATIVE COVALENT BOND

The coordinate covalent bond is also a special type of covalent bond, in which both electrons forming a bond are supplied by one atom only. The atom which supplies the pair of electrons for bond formation is known as ***"Donor"*** and the atom which accepts, is known as ***"Acceptor"***. An arrow (→)is commonly used to denote the bond, in which the head of an arrow pointing towards acceptor atom. The pair of electrons possessed by donor is called **"Lone-pair"** of electrons. This type of bond is also called dative covalent bond.

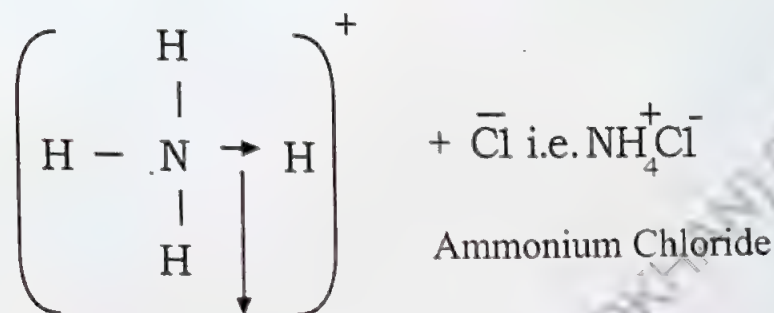
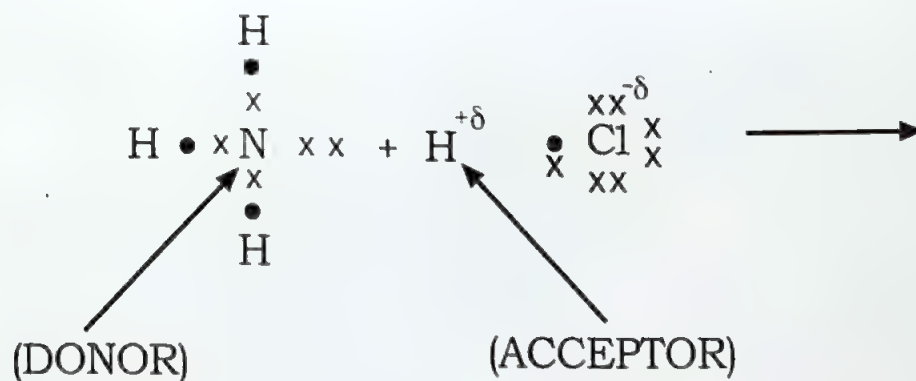
Examples :

In the formation of ammonium chloride, from ammonia (NH_3) and hydrogen chloride (HCl) molecules, the nitrogen (N) atom of ammonia (NH_3) acts as donor of an electron pair and hydrogen ion (H^+) of hydrogen chloride accept it to form (NH_4Cl) as follows.

$$H \bullet {}^{\times}\overset{\overset{H}{\bullet}}{\underset{\underset{H}{\bullet}}{N}} \times\times \; + \; H^{+\delta} \;\; \overset{\bullet}{\times}\,Cl^{-\delta} \longrightarrow$$

(DONOR) (ACCEPTOR)

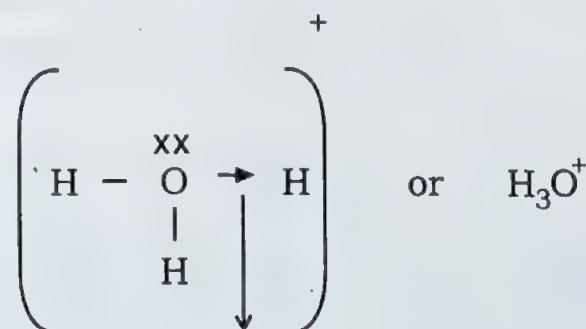
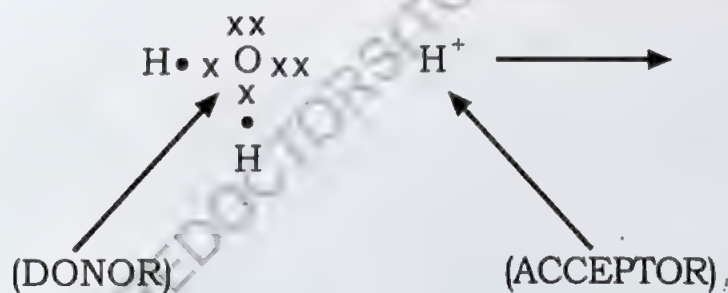
$$\left[\begin{array}{c} H \\ | \\ H - N \rightarrow H \\ | \\ H \end{array} \right]^{+} + \bar{Cl} \text{ i.e. } \overset{+}{NH_4}\overset{-}{Cl}$$

Ammonium Chloride

Co-ordinate covalent bond

Another example is provided by the formation of hydronium ion (H_3O^+) from water molecule and hydrogen ion (H^+),in which the oxygen atom of water acts as donor and hydrogen ion (H^+), as acceptor.

$$H \bullet \times \overset{\times\times}{O} \times\times \; H^{+} \longrightarrow$$

(DONOR) (ACCEPTOR)

$$\left[\begin{array}{c} \times\times \\ H - O \rightarrow H \\ | \\ H \end{array} \right]^{+} \text{ or } H_3O^{+}$$

Co-ordinate covalent bond

5.6.1 Comparison Between Covalent Bond and Co-ordiante Covalent Bond:

Covalent bond	Co-ordinate Covalent bond
1 The covalent bond is formed by the mutual sharing of electrons between atoms.	1. The co-ordinate covalent bond is formed by one sided sharing of electrons.
2. The covalent bond is formed between the similar or dissimilar atoms, when the electrons are mutually shared.	2. Co-ordinate covalent bond is formed between two unlike atoms, one having an electron pair available for sharing and other must accept that electron pair.
3. Covalent bond may be polar or non-polar	3. Co-ordinate covalent bond is always polar, because it is not formed between like atoms and is also known as co-ionic.
4. Covalent bond is associated with only covalent character because there is no electron transfer.	4. Co-ordinate covalent bond is associated with the ionic and covalent character because of partial transfer of electrons.
5. The shared pair of electrons in covalent bond are denoted by short lines e.g. for single pair (-) for double pair (=) and for triple pair of electrons (≡).	5. The shared pair of electrons in co-ordinate covalent bond is denoted by an arrow (→).
6. Covalent compounds are usually insoluble in water.	6. Where as co-ordinate covalent compounds are sparingly soluble in water.

5.7 METALLIC BONDING

We have noticed in covalent bonding, that non-metals have sufficient number of valence electrons to combine and form molecules by sharing of electron pairs with one another. For example two chlorine atoms share an electron pair in Chlorine (Cl_2) molecule.

However most of the metal atoms have less than four valence electrons (many metals have only one or two). These electrons are not confined to any particular atom, instead they move freely from one atom to another atom. Hence the atoms should be considered to be positively charged ions, therefore, metal is defined as: *A substance consisting of positively charged ions, fixed in a crystal lattice with negatively charge electrons moving freely through the crystal.* Therefore free electrons act as cohesive force which hold the atoms together and form a metallic bond.

Metallic bond is defined as the combination of electrostatic attraction between the electrons and the positive nuclei of atoms. X-rays analysis reveals that metal particles are held together in a lattice of closely packed spheres.

A survey of the observed properties of metals indicates, that the nature of force, holding the atoms together in a crystal, must be unusual. Metals are ductile(easily converted into thin wires) and malleable (easily bent or hammered into sheets).

They are good conductors of both heat and electricity (in solid or liquid states), they appear shiny and lustrous. Metals are solids at room temperature (except mercury (Hg)). They can be mixed with metals or non-metals to yield alloys with variable properties. All of these properties are because of peculiar bonding i.e. metallic bonding.

The strength of metallic bond varies considerably among different metals. Thus, it is much stronger in iron than in sodium or potassium.

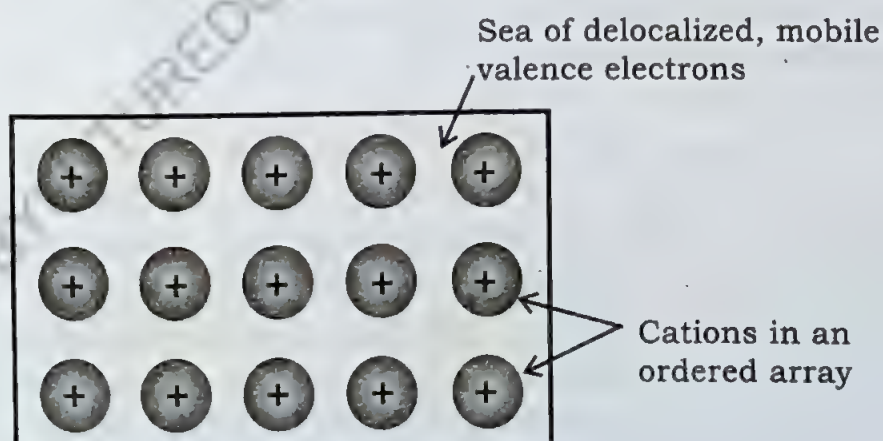



Fig. 5.8 Diagramatic representation of metallic bonding

5.7.1 Metallic Bonding in Sodium: (Na)

Sodium (Na) metal has one valence electron per atom. It crystallizes in a body centered cubic structure, in which each sodium is surrounded by eight nearest neighbours. These valence electrons are not confined to any particular atom, instead, they are free to move throughout the crystal, so the resulting bond is relatively weak, that is, why metals like sodium and potassium are soft and have relatively low melting points.

5.7.2 Metallic Bonding in Iron (Fe) and Copper (Cu):

The metals iron (Fe) and copper (Cu) are hard and have high melting points because, these metals have incomplete valence shells. Therefore, the atoms become covalently bonded to each other through their unfilled orbits. As a result strong covalent bonding between atoms extends throughout the crystals. This accounts for their hardness and high melting points.

5.7.3 Explanation of the Properties of Metals:

Since electrons in metals are free to move from one atom to the next. They are generally good conductors of electricity.

When metal is heated, the mobile electrons absorb heat energy and transfer to neighbouring electrons, this means that metals are good conductors of heat.

The mobile electrons readily absorb light falling upon them and move to higher energy levels. When they fall back to their original position they emits radiations. This causes the metallic lustre.

5.8 INTER-MOLECULAR FORCES (Vander Waals's Forces)

There are two types of forces present in molecules, that is:

(1) Intra-molecular forces and

(2) Inter-molecular forces.

Intra-molecular forces, hold atoms together in a molecule. For example, water (H_2O) molecule consists of two hydrogen atoms and one oxygen atom joined together in a specific way, that we call covalent bonds.

Inter-molecular forces, are the attractive forces between the neutral molecules, which hold them together at certain temperature. These attractive forces of neutral molecules between each other are called " Vander Waal's Forces", named after the Dutch Physicist (1837-1923), who first described them. There are three types of attractive forces between molecules (1) Dispersion forces (2) Dipole-dipole forces, induced dipole and (3) Hydrogen bonding.

Generally, inter molecular forces are much weaker than intra-molecular forces.

1. Dispersion Forces (London forces):

All particles, whether individual atoms, ions or molecules experience dispersion forces, which result from the motion of electrons around atom. For example consider atoms of noble gases e.g. He, Ne, Ar. etc. Let us examine the attractive forces in neon as an example. The distribution of ten electrons around the nucleus of neon is spherically symmetrical. But in case when two (Ne) atoms, come extremely close together. The electron clouds will repel each other. This polarizes each molecule and gives rise to an induced or temporary dipoles and as a result weak attractive forces called dispersion forces also called London forces after Fritz London (who first indentified them in 1930) are developed. The attraction is strong when particles are close together but rapidly weakens as they move apart.

Ne

Ne

Induced or temporary dipoles in neon molecules

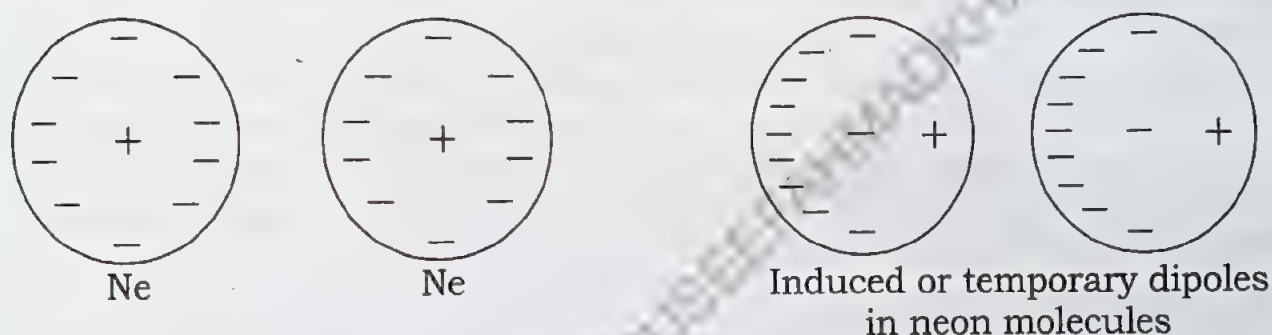



Fig. 5.9 Dispersion forces

Thus dispersion forces (London forces) are the weak attractive forces between temporarily polarized atoms (or molecules) caused by the varying positions of the electrons during their motion about the nuclei.

London forces are generally small as their energies are in the range of 1-10 KJ/mol.

2. Dipole-Dipole forces:

Dipole-dipole forces, are forces that act between polar molecules that possess dipole moments. *A Dipole-dipole force, is an attractive inter-molecular force resulting from the inter action of the positive end of one molecule with the negative end of other.*

Dipole-Dipole forces are generally stronger than dispersion forces.

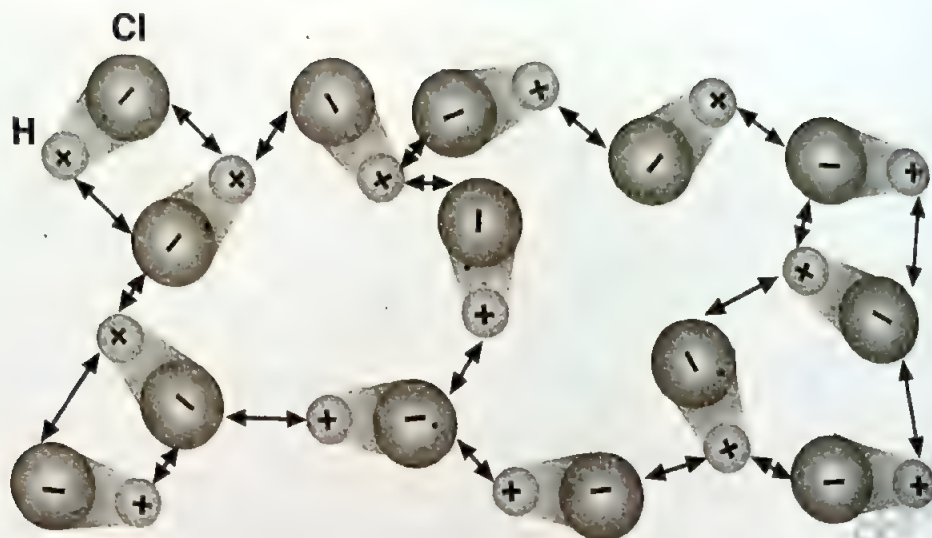



Fig. 5.10 Dipole - Dipole forces

Consider the (H-Cl) molecule, as chlorine has greater electronegativity than hydrogen, a partial negative charge on chlorine atom, and a partial positive charge on hydrogen atom is developed. The $(H^{+\delta}-Cl^{-\delta})$ has a permanent dipole moment. The electrostatic attraction of positive end of one (HCl) molecule for the negative end of another, constitutes an attractive forces in addition to dispersion forces. As a result polar (HCl) boils at (-85°C) but non-polar (F_2) boils at (-188°C) though both have nearly same molecular weights (36.46 a.m.u for HCl and 38.a.m.u for F_2).

3. Hydrogen Bonding:

When non-metal atoms of high electronegativity like those of F, O and N, are linked to hydrogen, their exist a force of attraction between positive hydrogen atom of one molecule and negative oxygen, nitrogen or fluorine atom of another. This force is so strong enough to cause two or more molecules to associate in larger clusters, as for example, $(H_2O)_x$ and $(NH_3)_x$.

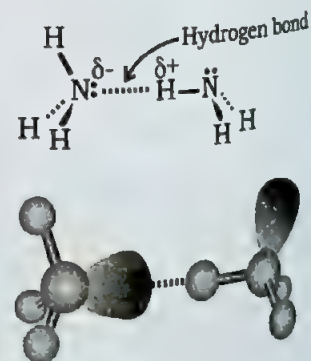
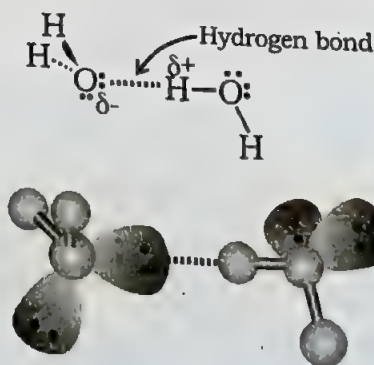
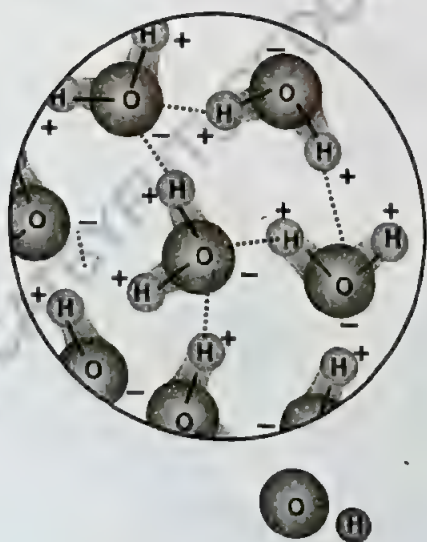



Fig. 5.11 Hydrogen bonding in water and ammonia

This attraction between positive hydrogen and negative oxygen or fluorine, is called hydrogen bond. This attraction or "hydrogen bond" can have about 5% to 10% of the strength of covalent bond. Hydrogen bond is denoted by dotted lines (-----).

Hydrogen bondings differ from the word "bond" since it is a force of attraction between positive hydrogen atom of one molecule and negative oxygen atom of another molecule. That is, it is an intermolecular force, not an intra-molecular force as in the common use of word bond. For this, hydrogen bonding is particularly strong type of dipole-dipole inter-action.

Hydrogen bondings have an important affect on the properties of water and ice. Hydrogen bonding is also very important in proteins and nucleic acids and therefore in life processes.

SUMMARY

1. Chemical bonds hold groups of atoms together to form molecules or solids. Bonding occurs when a group of atoms can lower its total energy by combining.
2. Bonds can be classified into two main types; ionic bond and covalent bond.
3. Electronegativity is defined as the relative ability of an atom in a molecule to attract the shared electrons in a bond to itself. The Electronegativity difference of the bonded atoms determines the polarity of that bond.
4. A covalent bond may be single bond, a double bond or a triple bond, depending on whether the two atoms share one, two or three pairs of electrons.
5. The co-ordinate covalent bond is formed only when an atom with an unshared pair of electrons in its valence shell donates a pair of electrons to another atom or ion that needs a pair of electrons to acquire a stable electronic configuration.
6. Covalent molecules are partially ionic, if they exist between two unlike atoms and their ionic character depends upon difference in electro-negativities of bonded atoms.
7. Metallic bond is the electrostatic attraction between positive ions and electrons of the atoms.
8. A hydrogen bond is a dipole-dipole attractive force that exist between two polar molecules, containing a hydrogen atom covalently bonded to an atom of F, O or N. Water is best example of hydrogen bonding.

EXERCISE

1. Fill in the blanks:

(i) .... covalent molecule is electrically neutral as well as symmetrical.

(ii) The power of an atom to attract the shared pair of electrons towards itself is called.........

(iii) .... compounds are usually made up of discrete units; with weak inter molecular forces.

(iv) NaCI is an .... compound.

(v) If electronegativity difference of bonded atoms is more then 1.7, the bond is............

(vi) The electrostatic attraction between positive ions and the electrons of the atoms is called .... bond.

(vii) The forces which hold atoms together in a molecule are called.......

(viii) The attraction between the positive hydrogen and negative F, O or N is called .... bonding.

(ix) CO_2 is a .... molecule.

(x) The atom which accepts a lone pair of electron is called...............

2. Tick the correct answer:

(i) The force which hold atoms together in a molecule or crystal is called:
(a) Ionic bond.
(b) Covalent bond.
(c) Co-ordinate covalent bond.
(d) Chemical bond.

(ii) The bond which is formed by the transfer of one or more electrons from one atom to another atom is called:
(a) Ionic bond.
(b) Covalent bond.
(c) Co-ordinate covalent bond.
(d) Chemical bond.

(iii) The bond which is formed by the mutual sharing of electrons between the atoms, is called:
(a) Ionic bond.
(b) Covalent bond.
(c) Co-ordinate covalent bond.
(d) Chemical bond.

(iv) The bond which is formed by one sided sharing of pair of electrons is called:

(a) Ionic bond. (b) Covalent bond.
(c) Co-ordinate covalent bond. (d) Chemical bond.

(v) The bond in MgO is:

(a) Electro-valent bond. (b) Covalent bond.
(c) Co-ordinate covalent bond. (d) Chemical bond.

(vi) The shared pair of electrons which links the atoms in a molecule is known as:

(a) Electro-valent bond. (b) Covalent bond.
(c) Co-ordinate covalent bond. (d) Chemical bond.

(vii) Double covalent bond is denoted by:

(a) Single short line. (b) Two short lines.
(c) Three short lines. (d) None of these.

(viii) The atom which supplies the pair of electrons for bond formation is known as:

(a) Acceptor. (b) Donor.
(c) Receiver. (d) None of these.

(ix) Co-ordinate covalent bond is always formed between the two:

(a) Like atoms. (b) Unlike atoms.
(c) Similar atom. (d) Like and unlike atoms.

(x) The shared pair of electrons in a co-ordinate covalent bond is denoted by:

(a) A single line. (b) Double line.
(c) An equal sign. (d) An arrow.

3. Write answer of the following questions:

(i) Define chemical bond? Discuss how atoms unite and change into molecules?

(ii) What are the valence electrons of an atom? How many valence electron does a nitrogen atom possess?

(iii) What happens to electrons, when elements combine?

(iv) What part of the atom is involved in the formation of chemical bond?

(v) Explain with examples? How elements are united by electro-valent bond?

(vi) What common properties are shown by ionic compounds?

(vii) What is meant by covalent bond? Write electronic formulas of any two covalent molecules? What is single, double and triple covalent bond?

(viii) Draw the electronic formulae for the following covalent molecules?
(a) H_2 (b) O_2 (c) N_2 (d) C_2H_2 (e) CO_2

(ix) Classify the following bonds as ionic or covalent. For those bonds that are covalent indicate whether they are polar or non-polar.
(a) H_2 (b) H-Cl (c) NaCl (d) $CaCO_3$ (e) $HC \equiv CH$ (f) O=O

(x) What are types of chemical bondings?

(xi) Account for the fact that some covalent bonds are polar while others are non-polar.

(xii) What is co-ordinate covalent bond? Explain with examples?

(xiii) Define the term covalent bond? How does a covalent bond differ from co-ordinate covalent bond?

(xiv) Explain electronegativity.

(xv) Explain in your own words Pauling (E.N.) table? Explain its usefulness in predicting the relative ionic and covalent character of a given compound?

(xvi) Give the characteristics of covalent compounds?

(xvii) What do you understand by ionic character of covalent bond? Under what conditions are the following formed
(a) Polar covalent bond.
(b) Non-polar covalent bond.
(c) Ionic bond.

(xviii) Define the term metal? And describe metallic bond?

(xix) Explain the following properties of metals?

(a) Lustre. (b) Conductivity.
(c) Malleability. (d) Ductility.

(xx) Why are some metals, such as sodium is soft, while other are hard?

(xxi) Explain the origin of dipole-dipole forces between the molecules? Give an example?

(xxii) What do you mean by dispersion forces? Why they are also called London forces?

(xxiii) What is hydrogen bonding? What type of forces, either intra-molecular or inter-molecular forces are present in hydrogen bonding?

CHAPTER-6

STATES OF MATTER

You will learn in this chapter about:

* Common states of matter.
* Kinetic description of states of matter.
* Solid state of matter.
* Liquid state of matter.
* Gaseous state of matter.
* Inter conversion of the three states.
* Diffusion in gases and liquids.
* Brownian movement.

6.1 COMMON STATES OF MATTER

Matter is any thing that has mass and occupies space. There are three common states of matter i.e solid, liquid and gas. Solid has definite shape as well as definite volume. Liquid has a definite volume but not definite shape and it takes the shape of vessel in which it is kept. Gas has neither definite shape nor definite volume, it takes the shape of vessel in which it is kept and occupies all the space available to it.

6.1.1. Kinetic Description of States of Matter

According to kinetic molecular theory, matter is composed of very tiny particles, which are called as molecules. These molecules are always in motion and they possess kinetic energy. The three states of matter depend upon the arrangements, motions and forces of attraction between these particles. The difference between solid, liquid and gas is explained as follows:

Solid State: In solid state molecules or particles are tightly packed with one another and they have only back and forth motion (Vibrational motion) about their fixed positions. Therefore particles in solid can not slip or slide over one another and thus they possess definite shape and definite volume.

Liquid State: In liquid state molecules are not tightly packed with one another. Their positions are not fixed and they can move in all direction, hence liquid does not have any definite shape. But in liquid state the kinetic energy of molecules is less than gaseous state. Therefore intermolecular attractive forces are more than gaseous state. Due to these intermolecular attractive forces liquid has a fixed volume.

Gaseous State: In gaseous state molecules are lying away from one another and they can move in all direction easily. The cohesive forces in a gas are negligible and the paritcles are free to move about in all directions. The kinetic energy possessed by these molecules are very high, therefore gases have neither definite shape nor definite volume.

6.1.2 Inter Conversion of Three States

Most of the important properties of liquids, solids and gases depend upon the interconversion of one state into another state. For example when ice is warmed, it melts into liquid state, water. When water is heated and boiled, it changes into vapour (gas) state, some solids upon heating instead of changing into liquid state are converted directly into gaseous (vapour) state. This process is called sublimation. The examples of such substances are iodine, ammonium chloride and napthaline. These chemical substances being in solid state, when heated are sublimed directly in vapour state. In general, each state may be converted into another state is explained as given below:

When a solid is heated, then the kinetic energy of particles increases and it becomes hot. If heating is continued then at certain temperature, the added energy becomes enough to overcome the attractive forces holding particles of solid in their fixed positions and it starts melting. The temperature at which solid starts melting is called melting point. At this point solid particles lose their fixed positions as well as their arrangements and thus solid is converted into liquid. At melting point the temperature of solid and liquid remains same untill all the solid is melted though heat. It means in this condition the rise in temperature is used in change of state from solid to liquid and this change is called **fusion**.

When liquid is heated, the kinetic energy of liquid molecules increases. Due to increase of kinetic energy certain molecules start escaping from

the surface of liquid and this escape of molecules is called ***evaporation***. At certain temperature the vapour pressure of liquid becomes equal to external pressure and at this temperature liquid starts boiling. This temperature is called **boiling point**. *Actually boiling point is the temperature at which the vapour pressure of a liquid becomes equal to external pressure applied on the liquid.* At this point bubbles of vapours are able to form within the interior of liquid and then rise to the surface, where they burst and release vapours and thus liquid is converted into gas. At boiling point the temperature of liquid remains same until all of the liquid is evaporated or vapourized though heat. It means in this condition the rise in temperature is used in the change of state from liquid to gas, and this change is called ***vapourization.***

6.2. DIFFUSION IN GASES AND LIQUIDS

6.2.1 Diffusion of Gases

It is observed that when a sample of a gas is set free in one part of the closed container, its molecules very quickly spread through out the container. In some cases we can smell a gas as it diffuses throughout the closed room. For example molecule of perfume spread throughout room or smell of H_2S gas spreads in laboratory etc.

The spreading of a substance through medium like air or liquid is called ***diffusion***. The rate of diffusion of a substance depends upon its molar mass or density.

A Scottish chemist, Thomas Graham in 1846 studied the rate of diffusion of different gases and formulated Graham's law. This law states that, the rate of diffusion of a gas is inversely proportional to the square root of its molar mass or density.

In other words, lighter gases can diffuse faster than heavier gases. For example hydrogen gas diffuses four times faster than oxygen gas at similar conditions.

6.2.2 Diffusion of Liquids

Liquid is intermediate between gaseous and solid states. Like gases liquid molecules are able to move and thus flow and diffuse. The rate of movement of liquid molecules is smaller than gases, hence they diffuse slower than gases. For example add two or three drops of blue ink in 200mls of water in a beaker. It is seen that blue colour of ink spreads

slowly in water and the whole water becomes bluish after some time. It means that diffusion also occurs in liquids but the rate of diffusion in liquids is slower than gases.

6.3 BROWNIAN MOVEMENT

This property was first of all observed by British Botanist, Robert Brown in 1827, during the movement of pollen grains in water by microscope.

If a drop of liquid with particles suspended in it, is observed under a microscope, it is seen that the particles are not at rest but they are moving in all directions in zig-zag motion. This motion is called as Brownian movement after the name, who observed it. *"A continuous, rapid, zig-zag motion of suspended particles through the medium is called Brownian motion."*

Example: Mix some powdered sulphur in water and stir it, after stiring filter the suspended sulphur. Some of the sulphur particles are very small and they can pass through the pores of filter paper into filterate. Now put a drop of this filterate on a slide and examine it under high powered microscope. It is observed that sulphur particles perform rapid random zig-zag motion through the medium and this motion is called Brownian motion.

SUMMARY

1. Any thing, that has a mass and occupies space, is called matter.
2. The common states of matter are solid, liquid and gas. Gas has neither definite shape nor volume, liquid has definite volume but no definite shape, whereas solid has definite shape as well as definite volume.
3. Liquid takes the shape of vessel in which it is kept, but does not occupy total volume available, whereas gas takes the shape of vessel as well as occupies all the space available to it.
4. According to kinetic molecular theory, matter is composed of particles, which are called molecules. Molecules are in motion and they possess kinetic energy.
5. The three states of matter depend upon arrangement, motion and forces of attraction between their particles.
6. In solids molecules are tightly packed with one another and they perform only translational motion. Due to this, molecules in solid neither slip nor slide over one another, hence shape and volume of solid is definite.

7. Interconversion of three states of matter are common and they may be seen by us, just like change of ice to water, evaporation of water and sublimation of iodine, ammonium chloride and napthaline.

8. During melting, solids are converted into liquids and in this process movement of particles overcome the intermolecular attractive forces.

9. Certain solids are directly converted into gaseous state during heating and this process is called sublimation.

10. In gaseous state, molecules are widely separated from one another and they move freely in all direction, hence the shape and volume of gas are not fixed.

11. When vapour pressure of a liquid becomes equal to external pressure, then it starts boiling and the temperature is called the boiling point. During boiling bubbles of vapours are able to form within the bulk of liquid and then rise to the surface, where they burst and release vapours. In other words in this way liquid is converted into gas.

12. At boiling point the temperature of liquid remains same until all of the liquid is vapourized to gas. This change is called as vapourization or evaporation.

EXERCISE

1. Fill in the blanks:

(i) There are ..... common states of matter.

(ii) Gas possesses neither definite shape nor definite ....

(iii) All types of matter is usually composed of smallest particles which are always in ....

(iv) The temperature at which liquid starts boiling is called .... Point.

(v) The .... is the intermediate state between solid and gas.

(vi) .... is the escape of molecules from the surface of liquid.

2. **Tick the correct answers:**

(i) The number of common states of matter is:
(a) 5 (b) 4
(c) 3 (d) None of them.

(ii) Neither definite shape nor volume is the property of
(a) Solid (b) Gas
(c) Liquid (d) None of them.

(iii) The temperature at which the vapour pressure of a liquid, becomes equal to its external pressure is called:
(a) Melting point (b) Boiling point
(c) Triple point (d) Freezing point.

(iv) The state of matter in which molecules are tightly packed and possess only transition motion is
(a) Gaseous state (b) Solid state
(c) Liquid state (d) None of them.

(v) The process in which molecules escape from the surface of liquid is called:
(a) Sublimation (b) Evaporation
(c) Boiling (d) Melting

(vi) The process in which solid directly changes to gas is called.
(a) Evaporation (b) Melting
(c) Sublimation (d) None of them

3. **Write answer of the following questions:**

(i) Define following terms:
(a) Solid (b) Fusion
(c) Evaporation (d) Sublimation

(ii) What do you mean by liquid? Describe the interchange of liquid to gas.

(iii) How solid is converted into liquid? Explain it.

(iv) What is diffusion? Explain on the basis of kinetic molecular theory.

(v) What is Brownian movement? Describe with suitable example.

CHAPTER-7

SOLUTION AND SUSPENSION

You will learn in this chapter about:

* Solution.
* Types of solution.
* Saturated, unsaturated and super saturated solutions.
* Factors affecting solubility.
* Crystallization.
* Strength of a solution.
* Suspension.
* Examples of suspensions in daily life.

7.1 SOLUTION

A soulution is defined as a homogeneous mixture of two or more substances. The relative amounts of the components can be gradually changed within certain limits. The component of solution present in smaller amount is called **solute**, whereas the component present in greater quantity is called **solvent**. For example in a 5% glucose solution in water, solute is glucose and solvent is water. When water is solvent the solution is called **aqueous solution.**

Types of Solution:

Since matter exists in three states i.e. solid, liquid and gas, on mixing they produce nine types of solutions, which are listed as following:

Table.

No.	Solute	Solvent	Examples
1.	Gas	Gas	Air (Mixture of 78% N_2, 21% O_2 and 1% other gases)
2.	Gas	Liquid	Carbonated soft drinks such as coca-cola, Bubble-up etc. NH_3gas in water, and air dissolved in water.
3.	Gas	Solid	H_2 gas adsorbed over palladium metal
4.	Liquid	Gas	Cloud (water vapours in air), steam.
5.	Liquid	Liquid	Alcohol in water, water in milk, milk in tea, vinegar (acetic acid in water)
6.	Liquid	Solid	Amalgam (e.g. Mercury in sodium), and water in jelly powder.
7.	Solid	Gas	Smoke (carbon particles in air),
8.	Solid	Liquid	Sugar in water, sea water
9.	Solid	Solid	Alloys such as brass (copper and zinc) bronze (copper and tin), steel (carbon and iron), glass.

7.2 FACTORS AFFECTING SOLUBILITY

Many factors affect the solubility of a solute in a solvent. These factors may be:

1. Temperature.
2. Pressure (For gases).
3. Nature of solute.
4. Nature of solvent

Solubility and Temperature:

Solubility of a solid in liquid or solubility of partially miscible liquids increases with increase in temperature. Thus solubility of sugar in water at 0°C is 179 g/100ml whereas at 100°C it is 487 g / 100 ml.

But the solubility of gases in a liquid decreases with the increase in temperature. For this reason when a glass of cold water is warmed, bubbles of air are seen on the inside of the glass.

Solubility and Pressure:

The solubilities of solids and liquids are not affected by pressure. But the solubility of a gas in a liquid is directly proportional to the pressure of gas. This is called **Henry's Law** i.e.

$$m \propto P \text{ or } m = KP$$

Where "m" is the amount of gas dissolved.

This effect is used in manufacture of bottled soft-drinks as coca-cola; 7-up etc. These are bottled under a CO_2 pressure slightly greater than 1 atm. When the bottles are opened, pressure decreases, so solubility of CO_2 also decreases, hence bubbles of CO_2 come out of solution.

Solubility and Nature of Solute and Solvent:

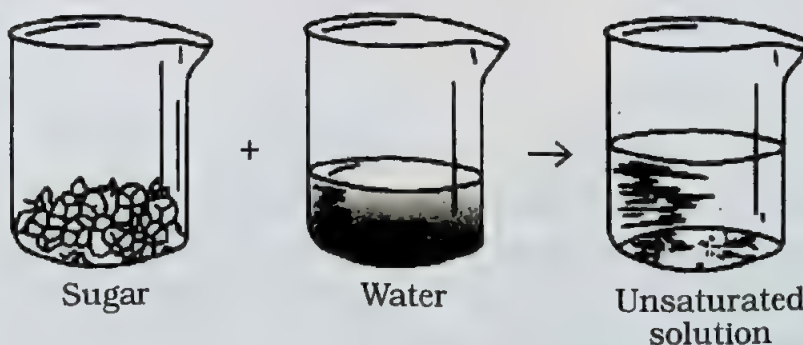
Solute and solvent may be polar (as H_2O; Alcohol) and non-polar (as benzene, carbontetrachloride). Polar and ionic solutes easily dissolve in polar solvents whereas non-polar solute easily dissolve in non-polar solvents. Hence it is said that **like dissolves like**. For example common salt (NaCl) being an ionic compound easily dissolves in polar solvent like water but is insoluble in non-polar solvent like benzene or petrol.

7.3 UNSATURATED, SATURATED AND SUPER SATURATED SOLUTIONS:

An unsaturated solution is that one which contains less solute than it has the capacity to dissolve. If more solute is added, at least some of it will dissolve.

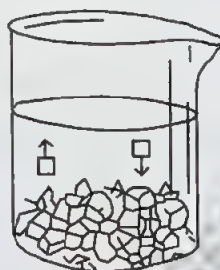
Preparation of unsaturated solution:

Take a few crystals of sugar and dissolve them in a glass of water. This results in an unsaturated solution, because the solution has a capacity to dissolve more crystals of sugar (solute) at a given temperature.

Sugar + Water → Unsaturated solution

A saturated solution is that one which contains maximum amount of solute in a given solvent at room temperature. No more solute can dissolve in it and there is a dynamic equilibrium between the undissolved and dissolved solute. This type of solution is called saturated solution.

The amount in grams of a solute required to be dissolved in 100 grams or 100 ml of a solvent to prepare a saturated solution at room temperature is called **solubility** of that solute, expressed in g/100 ml or g/100 g at that specific temperature.

Preparation of a Saturated Solution:

Take some water in a beaker. With constant stirring, add crystals of sugar till they do not dissolve and start to settle down. The solution thus obtained is saturated solution of sugar at room temperature.

A super saturated solution is that which contains greater amount of dissolved solute than that present in a saturated solution at the given temperature. Preparation of super saturated solution is required for crystallization. Crystallizaiton is simple and general technique for purification of impure compounds.

7.4 CRYSTALLIZATION

Crystals are homogeneous solids, having regular and definite geometrical shape with faces and sharp edges. Pure crystals of compounds have sharp melting points.

When a super saturated solution of a solid is prepared at high temperature and allowed to cool down, then at lower temperature it cannot hold more solute in dissolved state. Some of these dissolved solute particles come out of solution in solid form having regular and definite geometric shapes. They are called crystals. The process in which dissolved solute comes out of solution and forms crystals is called **crystallization**.

Following are two practical examples of super saturated solution preparation and crystallization.

Preparation of Crystals of Copper Sulphate (Blue Vitriol $CuSO_4.5H_2O$)

Prepare a saturated solution of copper sulphate in water at room temperature using a beaker. Heat the saturated solution and try to dissolve some more quantity of copper sulphate while stirring the solution with glass rod. Allow the super saturated solution of copper sulphate to cool down at room temperature. Upon cooling and standing, crystals of $CuSO_4.5H_2O$, i.e., blue vitriol will form. Filter out the crystals and observe the shape of crystals under a light microscope.

Preparation of Crystals of Potassium Nitrate (KNO_3)

Take 100 ml water in a beaker. Prepare a saturated solution of KNO_3 by dissolving 37g of solute at room temperature by means of stirring with glass rod. Now heat this saturated solution to 50°C and dissolve 20g of additional KNO_3 while stirring the solution. Filter the hot super saturated solution quickly and collect the filterate in another beaker. Cool the filterate to room temperature. Upon cooling crystals of KNO_3 are formed.

Filter out the crystals and observe their shape under light microscope.

Purification of Solids by Crystallization:

An impure solid generally contains two types of impurities. An insoluble impurity and a soluble impurity. Insoluble impurity is totally insoluble in the solvent used for crystallization even at boiling temperature. While the soluble impurity remains in soluble form at room temperature. Therefore, a compound containing such two types of impurities could be easily removed by means of crystallization technique.

For example, a 42g impure sample of KNO_3 contains a small quantity of sand and NaCl. To obtain pure crystalline KNO_3, we perform the crystallization technique as follows:

Take 50ml of water in a beaker and add the impure sample (40g) of KNO_3 to it while stirring with glass rod. Supply heat gently till the temperature of the solution is above 50°C. Stir the solution at this temperature till most of the solid is dissolved. Filter the hot solution and collect the filtrate in a beaker. Sand being insoluble in water will be removed and collected on the filter paper. Upon cooling of the filtrate, crystals of KNO_3 will start appearing. When no further crystals are formed, filter it again and collect the filtrate in a beaker. Purified crystals of KNO_3 are obtained on the filter paper. The filtrate will contain some quantity of the dissolved KNO_3 along with the NaCl, being a soluble impurity.

7.5 STRENGTHS OF A SOLUTION

The strength (concentration) of a solution means the mass or volume of the solute present in known amount of solvent or solution.

The following are the common methods of expressing the strength (concentration) of a solution.

1. Molarity (M)
2. Molality (m)
3. Mole fraction (x)
4. Percentage (%)
5. Normality (N). It is not used in these days.

1. Molarity (M):

It is defined as *the number of moles of solute dissolved in 1 liter ($1dm^3$) of a solution, it is denoted by (M).*

Thus, 1 mole of NaOH (i.e its gram formula mass) 40g dissolved in 1 litre ($1dm^3$) of solution is said to be 1 molar (1M) solution. If only half of the mole i.e 20g of NaOH is dissolved in one litre ($1dm^3$) of solution, the solution is said to be one-half molar (i.e. 0.5M)

The molarity of any solution is found by dividing the number of moles of solute by the number of liters of solution.

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Litres of solution or dm}^3 \text{ of solution}}$$

We know that:

(i) $$\text{Number of moles} = \frac{\text{Given mass of solute}}{\text{Gram formula mass of solute}}$$

(ii) $$\text{Litres of solution} = \frac{\text{mls. of solution}}{1000}$$

Thus,

$$\text{Molarity= (M)} = \frac{\text{Mass of solute}}{\text{Gram formula mass of solute}} \times \frac{1000 \text{ cm}^3 \text{ (ml)}}{\text{ml (cm}^3\text{) of solution}}$$

For example:

Calculate the molarity of a solution containing 4g of NaOH in 500 cm^3 (ml) of solution?

Solution:

$$M = \frac{\text{Mass of solute} \times 1000\ cm^3}{\text{Gram formula mass} \times cm^3 \text{ of solution}}$$

Data:

i)	Molarity (M)	=	?
ii)	Mass of solute NaOH	=	4g
iii)	Gram formula mass of solute	=	40g
iv)	cm^3 of solution	=	$500cm^3$

$$\therefore M = \frac{4g \times 1000\ cm^3}{40g \times 500cm^3} = \frac{1}{5} = 0.2M$$

Result: Molarity (M) = 0.2M

2. Molality (m):

It is defined as *the number of moles of solute dissolved per 1000g (1kg) of solvent, it is denoted by (m).*

Thus 1mole of Na_2CO_3 (i.e. its gram formula mass) 106g dissolved in 1000g of solvent is said to be 1 molal (1m) solution.If only half of the mole i.e. 53g of Na_2CO_3 is dissolved in 1000g of solvent, the solution is said to be one-half molal (i.e. 0.5m)

The molality of a solution is found by the following formula.

$$\text{Molality (m)} = \frac{\text{Mass of solute} \times 1000g}{\text{Gram formula mass} \times \text{grams of solvent}} \text{ or } \frac{\text{No of moles of solute}}{\text{Kg of the solvent}}$$

For example:

Calculate the molality of solution containing 5.3g of Na_2CO_3 in 500g of water.

Solution:

$$\text{Molality (m)} = \frac{\text{Mass of solute} \times 1000g}{\text{Gram formula mass of solute} \times \text{grams of solvent}}$$

Data:

1.	Molality = (m)	=	?
2.	Mass of solute i.e. Na_2CO_3	=	5.3g
3.	Gram formula mass of solute Na_2CO_3	=	106g
4.	Grams of solvent (water)	=	500g

$$\text{Molality} = m = \frac{\text{Mass of solute} \times 1000g}{\text{Gram formula mass of solute} \times \text{grams of solvent}}$$

$$m = \frac{5.3g \times 1000g}{106g \times 500g} = \frac{1}{10} = 0.1m$$

Result: Molality = 0.1m

3. Mole fraction (X):

Mole fraction (X) of any component in a solution is the number of moles of the component divided by total number of moles making up a solution. It is denoted by X.

$$\text{Mole fraction (X)} = \frac{\text{Moles of component}}{\text{Total number of moles making up solution}}$$

For example, a solution is prepared by dissolving 1 mole of ethyl alcohol $C_2H_5\text{-}OH$ in 3 moles of water (H_2O), where n_A and n_B represent the number of moles of ethyl alcohol and water respectively.

Then

$$\text{Mole fraction of ethyl alcohol} = X_A = \frac{n_A}{n_A + n_B} = \frac{1}{1+3} = \frac{1}{4} = 0.25$$

$$\text{Mole fraction of water} = X_B = \frac{n_B}{n_A + n_B} = \frac{3}{1+3} = \frac{3}{4} = 0.75$$

Result: Mole fraction of ethyl alcohol $X_A = 0.25$

Mole fraction of water $X_B = 0.75$

Note, that sum of the mole fractions is equal to 1.

Mole fraction of ethyl alcohol	=	0.25
Mole fraction of water	=	<u>0.75</u>
Sum of the mole fractions	=	1.00

> ***Remember***
>
> *The mole fraction is dimensionless quantity that expresses the ratio of the number of moles of one component to the number of moles of all components present. The sum of mole fractions of all components of a solution must equal 1*

4. Percentage (%):

This is based on mass (m) or volume (v) of components of solution. It is of four types.

i) Percentage in $\frac{M}{M}$ $\left(\frac{\text{mass}}{\text{mass}}\%\right)$

ii) Percentage in $\frac{M}{V}$ $\left(\frac{\text{mass}}{\text{vol}}\%\right)$

iii) Percentage in $\frac{V}{M}$ $\left(\frac{\text{vol}}{\text{mass}}\%\right)$

iv) Percentage in $\frac{V}{V}$ $\left(\frac{\text{vol}}{\text{vol}}\%\right)$

Examples:

(i) 5% (M/M) solution means solute 5g in 95gm solvent.
(ii) 10% (M/V) solution means solute 10g in solution $100cm^3$.
(iii) 5% (V/M) solution means solute $5cm^3$ in solution 100 g.
(iv) 15% (V/V) solution means solute 15 cm^3 in 85 cm^3 solvent.

Problem of %:

6.5g of NaCl are dissolved in 43.5 g water. Calculate the percent by mass of NaCl in this solution.

Solution:

$$\%\text{ by mass of solute NaCl} = \frac{\text{Mass of solute} \times 100}{\text{Mass of solute + Mass of solvent}}$$

$$= \frac{6.5 \times 100}{6.5 + 43.5} = \frac{6.5 \times 100}{50.0}$$

Percent by Mass of NaCl = 13.0 %

7.6 SUSPENSION

If sugar or salt is dissolved in water, the crystals dissolve into molecules or ions. The resulting homogenous mixture is called true solution. If a fine sand is stirred in water, the crystals do not dissolve, but even after several days some of the smallest particles remain suspended, such mixture is called a **suspension.**

Suspension is defined as *a hetrogenous mixture consists of visible particles, each of which contains many thousands or even millions of molecules, surrounded by molecules of liquid.*

Important to note

Homogenous mixtures can be classified according to size of the constituent particles, as solution, suspension or colloids. In solution the dispersed partilces are of molecular size (0.1-1nm). In suspension, the dispersed particles are much larger than molecules (>1000nm).

Between these two extreme type of dispersion, there is another type of homogenous mixture in which dispersed particles are larger than molecules (2.0-1000 nm) but not large enough to settle out. This intermediate type of mixture is called colloid.

Difference Between Solution and Suspension

Solution	Suspension
1. The size of particles is between 0.1 to 1nm.	1. The size of particles is larger than 1000 nm.
2. Particles cannot be seen with low power microscope.	2. Particles can be seen by low power microscope.
3. It is homogeneous.	3. It is heterogeneous.
4. Particles do not settle down.	4. Particles settle down.
5. It is transparent.	5. It is not transparent.
6. Components cannot be separated by filtration.	6. components can be separated by filtration.

Examples of Suspensions in Daily Life:

1) **Smoke:** A suspension of the particles of carbon in a gas or air.
2) **Mud (slime):** A suspension of fine particles of solid in small quantity of liquid.
3) **Foam (froth):** A suspensionof fine particles of a gas in a liquid.
4) **Emulsion:** A suspension of droplets of one liquid into another in which it is not soluble.

SUMMARY

1. A homogenous mixture of two or more substances with uniform composition is called solution. A solution has two components, solute and solvent.

   A solute is the substance which is dissolved and is in smaller quantity while a solvent is the substance in which a solute is dissolved and is in larger quantity.
2. There are nine types of solutions depending upon the nature of solute and solvent. If solvent is liquid then the solution is considered as true solution.

3. A solution in which the amount of the solute is less than it has the capacity to dissolve in large quantity of solvent is called unsaturated solution.

   The solution which contains maximum amount of solute in a given solvent at a specific temperature and no more solute dissolves in it is called saturated solution. In a saturated solution there is a dynamic equilibrium between dissolved and undissolved solute.

4. A super saturated solution contains greater amount of dissolved solute than that are peresent in a saturated solution. It is obtained by dissolving solute in saturated solution on heating. More solute would dissolve on heating saturated solution.

5. The solubility is defined as the amount of solute in gram dissolved at a given temperature in 100 gram of the solvent. Many factors affect the solubility of a solute in a solvent which are temperature, pressure (for gases), nature of solute and solvent. Generally the solubility of a solute increases with the rise in temperature except gases whose solubility decreases with the rise in temperature.

   The solubility of gases in a solvent is generally less, but the solubility of gases increases with increase in pressure and is directly proportional to pressure applied. It is governed by Henry's Law.

   Nature of solute and solvent also affects on the solubility. It is well known fact that **like dissolves like.**

6. Crystallization is the process in which dissolved solute comes out of the solution and forms crystals. Crystals are homogenous solids bounded by plane faces, having sharp edges, regular and definite shapes. They have sharp melting points. Crystallization may occur by the process of evaporation or by cooling heated solution.

7. In a solution the amount of solute dissolved in a given quantity of solvent is known by its concentration. The solution that contains less amount of solute as compared to the amount of solvent is known as "dilute solution" and the solution that contains greater amount of solute as compared to the amount of the solvent is known as "concentrated solution".

8. The different units in which the concentrations of the solutions are expressed are molarity, molality, mole fractions, percent concentrations.

   Molarity is defined as the number of moles of the solute, dissolved per dm^3 or litre of the solution. It is denoted by M.

Molality is defined as the number of moles of the solute, dissolved in per kg or 1000 gram of the solvent. It is denoted by m.

Mole fraction is defined as the ratio of moles of solute or solvent to the total moles of the solution. It is denoted by X.

For example

$$X_{solute} = \frac{\text{No. of moles of solute}}{\text{Total moles of solute and solvent in the solution}}$$

Percent concentration is based on mass (M) and volume (V) of the components solute and solvent in the solution. There are four different ways in which percent concentration can be expressed.

$\% \frac{M}{M}$; $\% \frac{M}{V}$; $\% \frac{V}{M}$; and $\% \frac{V}{V}$

9. A suspension is a mixture in which particles are of one or more substances having size larger than 1000 nanometer (n.m) (1000 n.m or 1000×10^{-9}m). In suspension the solute particles are not soluble in solvent and can be removed by filteration because the solute particles are big enough not to pass from the pores of filter papers. Suspension is heterogenous.

EXERCISE

1. **Fill in the blanks:**
   (i) Solution is a .... of two or more substances.
   (ii) The most common solvent in nature is ....
   (iii) An .... is the solution when the liquid solvent is water.
   (iv) $10\% \frac{M}{M}$ solution contains 10 gram solute, dissolved in ......g solvent.
   (v) .... is the symbol for the concentration unit of molarity.

2. **Tick the correct answer:**
   (i) The suspended particles in suspensions are generally of the size.
   (a) 10nm (b) 100nm
   (c) 1200nm (d) 1nm

(ii) The sum of the mole fractions of solute and solvent is equal to:
(a) 5 (b) 2
(c) 0 (d) 1

(iii) Solubility is defined as the amount of solute in gram at a given temperature, dissolved in .... of the solvent.
(a) 20g (b) 100g
(c) 10g (d) 2000g

(iv) The process in which a solid directly changes to vapours is known as:
(a) Sublimation (b) Evaporation
(c) Diffusion (d) Fusion

(v) The solubility of a gas .... with the rise in temperature.
(a) Increase (b) Decrease

3. Tick true or false:

(i) The process of converting a solid into liquid at its melting point is called fusion.
(ii) A suspension is a homogenous mixture of two or more substances.
(iii) The solution that contains maximum amount of solute in a given solvent at specific temperature is a saturated solution.
(iv) Crystals have irregular geometrical shape.
(v) Smoke is a suspension of carbon particles in air.

4. Write answer of the following questions:

(i) Define the following terms:
(a) Solute (b) Solvent
(c) Solubility (d) Crystallization

(ii) Name the solute and solvent in the following solutions:
(a) Syrup (b) Haze (Dust in air)
(c) Butter (Water in fat) (d) Fog
(e) Jellies (Water in fruit pulp) (f) Smoke
(g) Sodium amalgam (h) Cheese (Water in fat)
(i) Foam (Water in air) (j) Mist

(iii) Discuss the factors affecting the solubility.

(iv) Explain Why?

(a) Common salt dissolves in water but not in petrol.

(b) Cold drinks are bottled under a CO_2 pressure greater than 1 atmosphare.

(c) 100 ml solution of KNO_3 can not hold more than 37gm of KNO_3 in dissolved state.

(v) Calculate molarity of solution containing 16 gm glucose per 300 ml solution.

(vi) Find the mass of sucrose (Molecular Mass=342) required to be dissolved per 600cm^3 solution to prepare a semi-molar solution.

(vii) 5.3 gm Na_2CO_3 was dissolved in 800gm water, calculate the molality of solution.

(viii) It is desired to prepare 3 molal solution of NaOH. How much mass of it must be dissolved in 1500gm water.

(ix) Differentiate between (a) Saturated and unsaturated solution, (b) Solution and suspension.

CHAPTER-8

ELECTROCHEMISTRY

You will learn in this chapter about:

* Electrolytes and non-electrolytes.
* Electrolysis.
* Electrolysis of molten NaCl.
* Electrolysis of water.
* Faraday's laws of electrolysis.
* Electrochemical equivalent.
* Uses of electrolysis.
* Electroplating.
* Electrochemical cells.
* Batteries.

INTRODUCTION

Chemical energy is associated with chemical reactions and the electrical energy is obtained from electricity. Thus chemical and electrical energies are the two forms of energy which are interconvertible.

The branch of chemistry that deals with the relationship between electricity and chemical reactions. It deals with the conversion of electrical energy into chemical energy and chemical energy into electrical energy is defined as electrochemistry.

8.1 ELECTROLYTES AND NON-ELECTROLYTES

The chemical compound which conducts electricity in molten condition or through its aqueous solution with chemical change is called an electrolyte.

All acids, bases and salts are electrolytes (in aqueous solutions or fused state). Such as Hydrochloric acid (HCl), Sulphuric acid (H_2SO_4), Sodium hydroxide (NaOH), Sodium chloride (NaCl), Copper sulphate ($CuSO_4$) etc, are electrolytes.

You cannot electrolyse something unless it conducts electricity. But all the conductors are not electrolytes,You can pass electric current through a metal for years but it will not decompose. Metals are good conductors but are not electrolytes. Metals are used as electrodes through which electric current enters and leaves the electrolytes.

All electrolytes are ionic compounds or polar compounds like acids and bases. Salts are ionic compounds and are solids. These solids do not conduct electricity because in the solid state the ions are very tightly packed or held together showing no movement of the ions. They cannot move. However when an ionic solid is melted or dissolved in water, its ions become free to move. Thus ionic compounds in molten states or in aqueous solution, conduct electricity. This conduction is due to the free movement of ions.

Non - Electrolytes:

The chemical compounds which do not conduct electricity in molten or in aqueous solutions are called non-electrolytes; such as Sugar, Petrol, Benzene etc.

8.2 ELECTROLYSIS OR ELECTROLYTIC CONDUCTION

A chemical process or reaction which does not have the capability to take place spontaneously, can be made to take place by passing electric current. If appropriate quantity of current is passed under suitable condition through an electrolytic solution, the reactions occurs non-spontaneously i.e. under stress.

The cell used for the reaction to occur by passing electric current is called electrolytic cell. Thus electrolytic cell is a device which is used to convert electrical energy into chemical energy (a non-spontaneous chemical reaction takes place by the loss and gain of electrons), for example a device for electroplating proceses or electrolysis of water in the presence of an acid. The process is called electrolysis or electrolytic conduction.

Electrolysis may be defined as a process in which movements of the ions take place towards their respective electrodes to undergo chemical changes under the influence of an applied electric field.

1. Electrolysis of Molten Sodium Chloride:

Sodium chloride, (Salt) does not conduct electricity in the solid state. To make it conduct electricity, either fuse or melt the salt or dissolve it in water. So there are two ways of doing electrolysis of sodium chloride. We are doing electrolysis of molten sodium chloride. Sodium chloride melts at 800°C. It is easy to predict the result of electrolysing a molten electrolyte. The salt just splits into two parts, sodium (metal) positive ion (Na^+) and chloride (non-metal) negative ion (Cl^-). We can write the ionization reaction as :

$$NaCl(s) \xrightleftharpoons{\text{fused}} Na^+(l) + Cl^-(l)$$

In solid sodium chloride, the ions are held together tightly in a regular lattice arrangements. In solid the ions cannot move, but when sodium chloride is melted, the ions are freed from their lattice and they can move freely to conduct electricity.

Some fused NaCl is taken in an electrolytic cell i.e. in a glass vessel. Two platinum rods (electrodes) are dipped into the fused salt (NaCl). The two electrodes are connected to a source of electricity i.e. to a battery outside the cell by wires. The electrode connected to the negative terminal of the battery is called cathode and the electrode connected to the positive terminal of the battery is called anode.

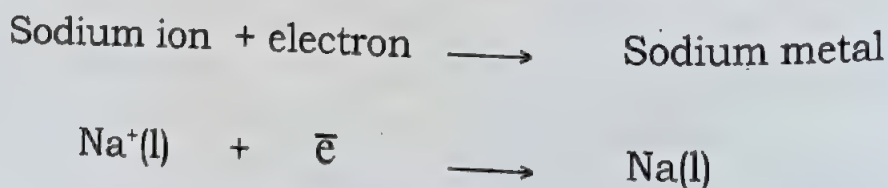
When an electrical potential is passed through the molten sodium chloride salt, electrolysis starts. The positive ions i.e. cations (Na^+) are attracted towards cathode and the negative ions i.e. anions (Cl^-) are attracted towards anode. At the two electrodes chemical reactions take place.

At Cathode:

Sodium ions (Na^+) which are positively charged,move towards cathode and gain electrons to get neutralized. Thus Na^+ ions are discharged at cathode and form neutral molten sodium metal.

We can represent the happening at cathode as :

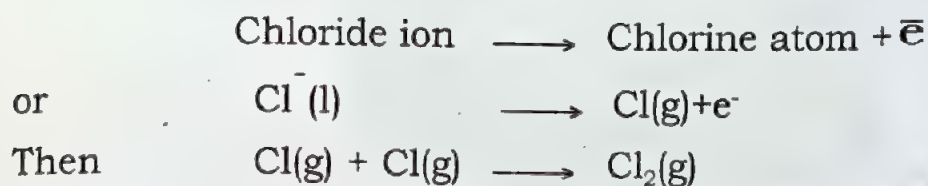
$$\text{Sodium ion} + \text{electron} \longrightarrow \text{Sodium metal}$$

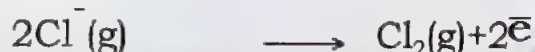
$$Na^+(l) + \bar{e} \longrightarrow Na(l)$$

At Anode:

Chloride ions (Cl^-) which are negatively charged move towards anode; Cl^- ions have one electron in excess. The anode has a shortage of electrons and is positively charged. When Cl^- ions arrive at anode, the anode attracts their electrons and Cl^- ions lose their extra electrons at anode, forming neutral chlorine (Cl) atoms but we know Cl atoms do not exist in the free state. They combine to form Cl_2 molecules. So Cl^- ions are discharged at anode to liberate Cl_2 gas by the loss of electrons at anode.

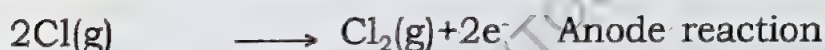
$$\text{Chloride ion} \longrightarrow \text{Chlorine atom} + \bar{e}$$

or $$Cl^-(l) \longrightarrow Cl(g) + e^-$$

Then $$Cl(g) + Cl(g) \longrightarrow Cl_2(g)$$

Over all it amounts to

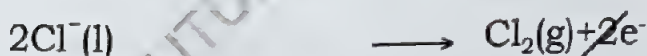
$$2Cl^-(g) \longrightarrow Cl_2(g) + 2\bar{e}$$

The over all reaction can be described as:

$$2(Na^+(l) + e^- \longrightarrow Na(l)) \qquad \text{Cathode reaction}$$

$$2Cl(g) \longrightarrow Cl_2(g) + 2e^- \qquad \text{Anode reaction}$$

On adding the two reactions, the loss and gain of electrons cancel and we get final result.

$$2Na^+(l) + \cancel{2e^-} \longrightarrow 2Na(l)$$

$$2Cl^-(l) \longrightarrow Cl_2(g) + \cancel{2e^-}$$

$$2Na^+(l) + 2Cl^-(l) \longrightarrow 2Na(l) + Cl_2(g)$$

The electrolysis of molten sodium chloride shows that we get sodium metal at cathode and Cl_2 gas is liberated at anode. The following fig 8.1 describes the electrolysis of molten NaCl.

$$2Cl^-(l) \longrightarrow Cl_2(g) + 2e^- \qquad 2Na^+(l) + 2e^- \longrightarrow 2Na(l)$$

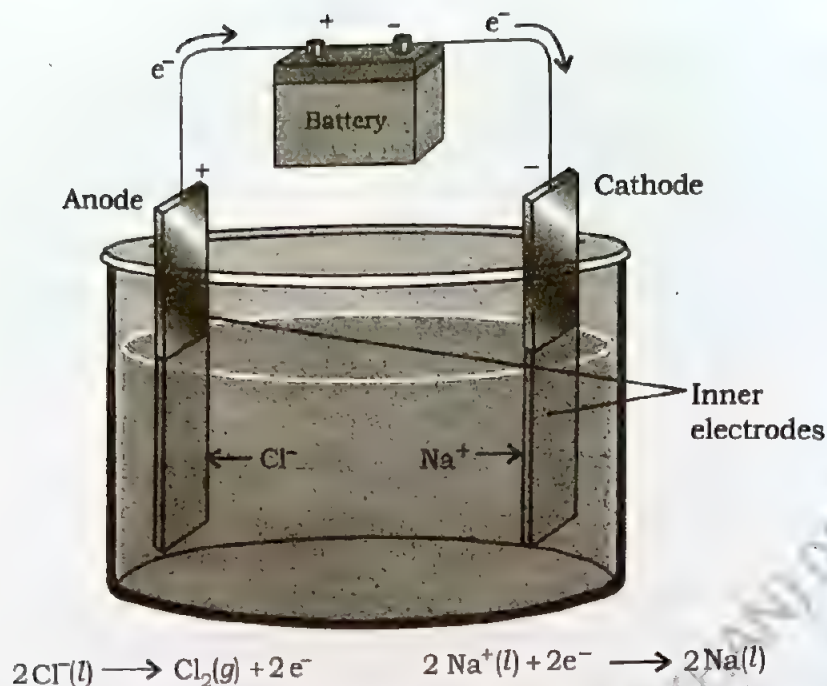
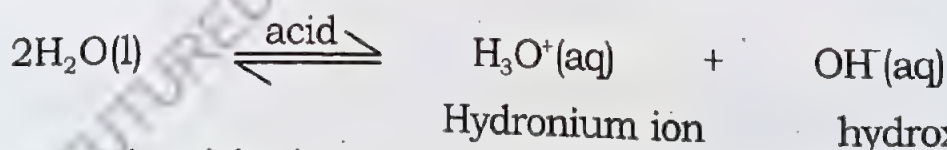



Fig. 8.1 Electrolysis of molten NaCl

2. Electrolysis of Water:

Pure water on its own hardly conducts electricity because it ionizes feebly into ions. But when few drops of acid or base or very small quantity of a salt is added in water, it makes water to conduct electricity and then water can be electrolysed to produce hydrogen gas at cathode and oxygen gas at anode.

Consider an electrolytic cell containing acidulated water. Two electrodes are dipped in the acidulated water. Water in the presence of few drops of acid ionizes as:

$$2H_2O(l) \xrightleftharpoons{acid} \underset{\text{Hydronium ion}}{H_3O^+(aq)} + \underset{\text{hydroxide ion}}{OH^-(aq)}$$

When electricity is passed from the battery, the positive ions (H_3O^+) move towards cathode. They gain electrons and get neutralized to liberate H_2 gas.

At Cathode:

$$2H_3O^+(aq) + 2e^- \longrightarrow H_2(g) + 2H_2O(l)$$

Hydroxide ion (OH^-) move towards anode. OH^- ions lose electrons at anode and get neutralized to liberate O_2 gas.

At Anode:

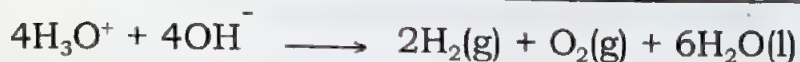
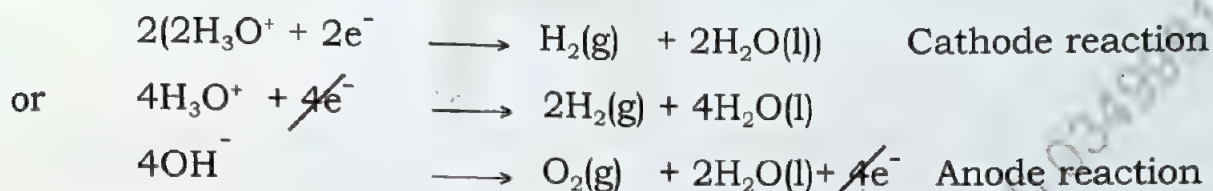
$$4OH^-_{(aq)} \longrightarrow O_2(g) + 2H_2O_{(l)} + 4e^-$$

The over all reaction should be balanced according to gain and loss of electrons at the two electrodes so the cathode reaction is multiplied by 2 and then added in the anode reaction.

$$2(2H_3O^+ + 2e^- \longrightarrow H_2(g) + 2H_2O(l)) \quad \text{Cathode reaction}$$

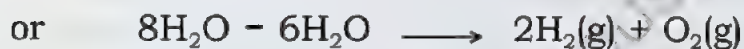
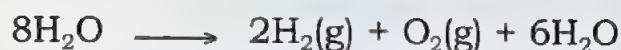
or $$4H_3O^+ + \cancel{4e^-} \longrightarrow 2H_2(g) + 4H_2O(l)$$

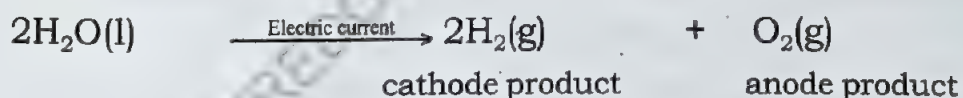
$$4OH^- \longrightarrow O_2(g) + 2H_2O(l) + \cancel{4e^-} \quad \text{Anode reaction}$$

$$4H_3O^+ + 4OH^- \longrightarrow 2H_2(g) + O_2(g) + 6H_2O(l)$$

$4H_3O^+$ and $4OH^-$ together will be equivalent $8H_2O$, So

$$8H_2O \longrightarrow 2H_2(g) + O_2(g) + 6H_2O$$

or $$8H_2O - 6H_2O \longrightarrow 2H_2(g) + O_2(g)$$

or $$2H_2O \longrightarrow 2H_2(g) + O_2(g)$$

For simplicity of the electrolysis of water, we write the equation as:

$$2H_2O(l) \xrightarrow{\text{Electric current}} \underset{\text{cathode product}}{2H_2(g)} + \underset{\text{anode product}}{O_2(g)}$$

It is observed that on electrolysis of water, we get two volumes of hydrogen gas per each volume of oxygen gas. Humphry Davy who first did the electrolysis of water, confirmed the formula of water is H_2O (2:1 ratio of Hydrogen and Oxygen).

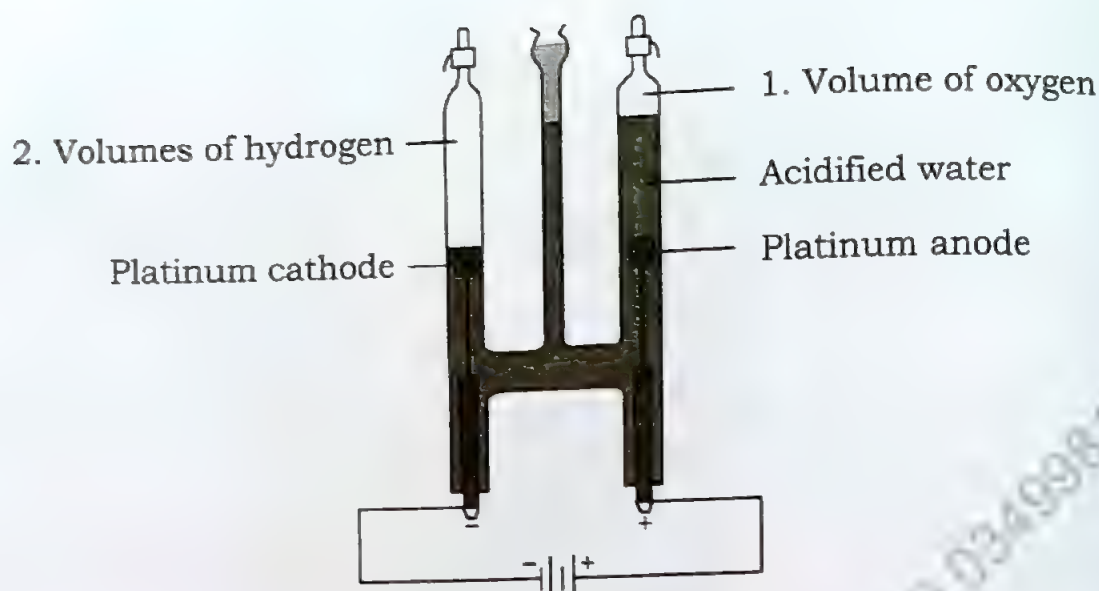



Fig. 8.2 Electrolysis of water

8.3 FARADAY'S LAWS OF ELECTROLYSIS

Michael Faraday in 1833, discovered the quantitative laws governing the process of electrolysis, which are known as Faraday's laws of electrolysis. There are two laws of electrolysis put forward by Faraday.

1. Faraday's First Law of Electrolysis:

It states that the amount of any substance deposited or liberated at an electrode during electrolysis is directly proportional to the quantity of current passed through the electrolyte.

Explanation:

If 'w' be the weight or amount of a substance deposited or liberated, and 'A' ampere of current is passed for 't' seconds, then according to the law:

$$w \propto A \times t$$

$$\boxed{\text{or} \quad w = Z A t}$$

Where Z is a constant, known as "electro chemical equivalent" of the substance. If one ampere of current is passed for one second; then w = Z. This means when one ampere of current is passed for one second; then the weight or amount of the substance deposited or liberated is exactly equal to its electrochemical equivalent. The current in ampere, multiplied by the time in second is known as coulomb which is the electric charge.

Ampere (A) x time (s) = coulomb (C)

Electrochemical Equivalent:

Electrochemical equivalent of a substance may be defined as the weight (amount) of the substance deposited or liberated, when one coulomb of electric charge is passed through an electrolyte. It is denoted by Z and in S.I unit it is expressed in Kg / coulomb. Each element has its own electrochemical equivalent.

Example - 1

A current of 0.5 ampere was passed through a solution of $CuSO_4$ for one hour. Calculate the mass of copper metal deposited on the cathode. Electrochemical equivalent of Cu=0.000329g/C = 3.29×10^{-4} g/C

Or 3.294×10^{-7} Kg/C.

Solution:

Data:

1. Current in ampere (A) = 0.5
2. Time in second (1 hr) = 1 x 60 x 60 = 3600 s
3. Z for Cu metal = 3.294×10^{-4} g/C = 3.294×10^{-7}Kg/C

Formula

$w = Z \times A \times t$

$= 3.294 \times 10^{-7} \times 0.5 \times 3600$

$= 5.929 \times 10^{-4}$ Kg/C

Mass of copper metal deposited = 5.929×10^{-4} Kg/C or 0.5929 g/C

Example - 2

A current of 10 amperes was passed for 15 minutes in a solution of silver nitrate ($AgNO_3$). The mass of silver deposited was found to be 1.0062×10^{-2} Kg. Calculate the electrochemical equivalent (Z) of Ag metal.

Solution :

Data :

1. Current in ampere = 10
2. Time in seconds (15 minutes) = 15 x 60 = 900 s
3. Mass of Ag metal deposited (w) = 1.0062×10^{-2} Kg

Formula

$w = Z A t$

or $\quad Z = \dfrac{w}{At} = \dfrac{1.0062 \times 10^{-2}\,Kg}{10A \times 900\,s\,C}$

$$= \frac{1.0062 \times 10^{-2}}{9 \times 10^{-3}} = \frac{1.0062 \times 10^{-2} \times 10^{-3}}{9}$$

$$= 0.1118 \times 10^{-5}\,Kg/C$$

$$= 1.118 \times 10^{-6}\,Kg/C$$

$$= 0.00118 \quad g/C$$

2. Faraday's Second Law Of Electrolysis:

It states that the masses of different substances deposited or liberated, when same quantity of current is passed through different electrolytes, connected in series are proportional to their chemical equivalent masses.

Consider three different electrolytes, $AgNO_3$, $CuSO_4$ and Al $(NO_3)_3$ solutions, connected in series. Same quantity of current is passed through them, then the masses of Ag, Cu and Al, deposited on their respective electrodes would directly proportional to their equivalent masses.

According to Faraday if exactly 96,500 coulombs of electric charge is passed then the mass of Ag deposited would be equal to 108g (108/1), that of copper is 31.75g (63.5/2) and Al is 9g (27/3) which are their equivalent masses respectively.

$$\text{Equivalent mass of an element} = \frac{\text{Atomic mass of the element}}{\text{Valency of the element}}$$

The current of 96,500 coulombs is called as one Faraday (F) charge after the name of Faraday. Thus Faraday is defined as the quantity of charge which deposits or liberates exactly one gram equivalent of a substance.

IF = 96500 Coulombs

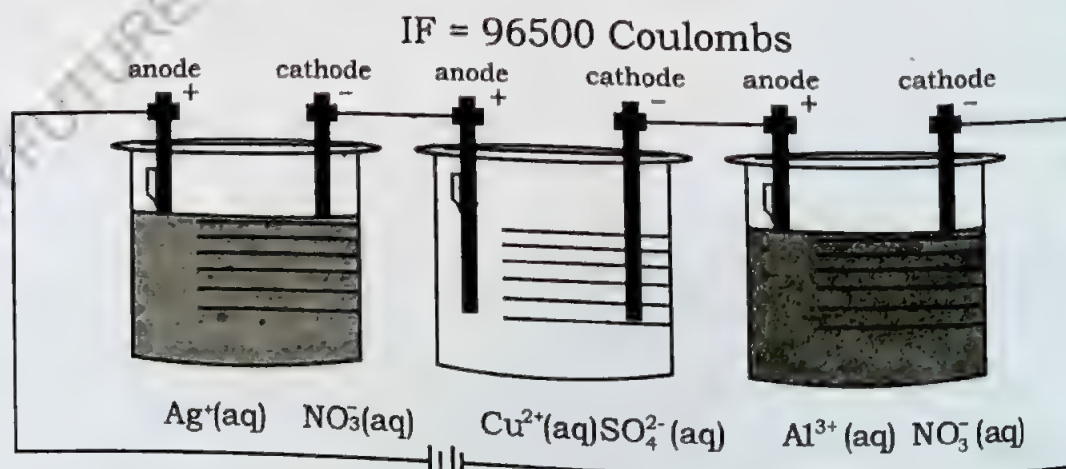



Fig.8.3 Electrolytic cells arranged in series

Relationship Between Equivalent Mass And Electrochemical Equivalent:

Since 96,500 C (IF) electric charge is required to liberate one gram equivalent mass of a substance, so it is clear that the gram equivalent mass of a substance is 96,500 times greater than its electrochemical equivalent.

Gram Eq.mass = 96,500 x E.C.E (Z)

In other words if e is the gram equivalent mass and Z is the E.C.E, then we can write it as :

$e = 96{,}500 \times Z$ or $e = F \times Z$.

UNITS OF CURRENT AND ELECTRIC CHARGE

1. **Ampere :**

It is the basic unit of current in the international system (S.I). It may be defined as the current when passed through a circuit for one second, can liberate 0.001118 g or 1.118×10^{-6} Kg of Ag from silver nitrate solution.

2. **Coulomb :**

It is the SI unit of electric charge and is defined as the quantity of charge when one ampere of current is passed for one second.

C	=	ampere x time (s)
1 coulomb	=	1 A x 1 s

A coulomb is equivalent to ampere multiplied by seconds. For example, a current of 0.5 ampere flowing for 80 seconds, gives a charge $0.5 \times 80 = 40$ coulombs.

If we are given the current and the time of electrolysis, we can calculate the amount of the substance produced at an electrode. On the other hand if we are given the amount of the substance produced at an electrode and the length of the time of electrolysis, we can determine the current or electrical charge.

Example-1:

When an aqueous solution of copper sulphate is electrolysed, copper metal is deposited at the cathode.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s) \text{ -- at cathode}$$

If a constant current was passed for 5 hours and 404 mg of Cu was deposited, calculate the current passed through $CuSO_4$.

Solution:

Amount of Cu deposited = 404mg = 0.404g.

Gain of $2e^-$ means 2F electric charge

Atomic mass of Cu = 63.5

According to cathode reaction.

63.5 g of Cu is deposited by 2 F electric charge

$\therefore$ 0.404g of Cu is deposited by $\frac{2}{63.5} \times 0.404$

$= 0.0127$ F

We know

$$1F = 96{,}500 \text{ coulomb}$$

$$\therefore 0.0127F = 0.0127 \times 96{,}500 = 1225.6C$$

$\therefore$ Coulomb = Ampere x t sec. (time = 5hours)

$$\therefore \text{Ampere} = \frac{C}{ts} = \frac{1225.6}{5 \times 60 \times 60} = 0.0680$$

$$= 6.80 \times 10^{-2} \text{ ampere}$$

Example-2:

How many grams of oxygen gas is liberated by the electrolysis of water after passing 0.0565 ampere for 185 second.

Solution:

Equation: $2H_2O \longrightarrow O_2(g) + 4H^+ + 4e^-$

According to equation (four) Faraday is required to liberate 32 g of O_2

Data:

Current in ampere = 0.0565

Time in second = 185

Coulomb = ampere x time (s)

= 0.0565 x185 = 10.45 C.

$$F = \frac{C}{96{,}500} = \frac{10.45}{96{,}500} = 0.000108 \text{ F}$$

Now

4 F electric charge liberates 32 g O_2

$\therefore$ 0.000108 electric charge liberates $\frac{32}{4} \times 0.000108$

$= 0.000864$ g

$= 8.64 \times 10^{-4}$ g O_2

8.4 USES OF ELECTROLYSIS

Electrolysis is an important process, used for the extraction of certain metals from their ores. It is also used in electroplating. Electroplating means to coat one metal on other metal through the process of electrolysis in order to protect the baser metals from corrosion and to make them more attractive.

In 1852, the price of Aluminium metal was very high and it was considered as an expensive oddity, because of its silvery - white shining property. However, by 1890, the price of Aluminium metal had fallen very low to a fraction of the price of silver. What had happened ?

Aluminium metal is a reactive metal which makes it difficult to extract it from its ore. Before 1886 the only way to get it, was by heating its salt Aluminium chloride with sodium metal. Sodium itself was expensive which made it even more expensive.

$$AlCl_3(s) + 3Na(s) \xrightarrow{heat} Al(s) + 3NaCl(s)$$

Then in 1886 a new process was developed which involved the electrolysing a molten Aluminum compound. The same basic method of electrolysis is still used for the extraction of Al-metal. Now-a-days Al is extracted by the electrolysis of Alumina (Al_2O_3) which is obtained from the chief ore of Al, bauxite (Al_2O_3 nH_2O). Due the process of electrolysis, the price of Aluminum quickly falls. Today Aluminum is so cheap that we can use disposable plates of Aluminum for eating and throw them away afterwords.

Many metals are extracted by the electrolysis of their molten compounds, usually their chlorides, because the chlorides of most of the metals have low melting points. For example sodium metal is extracted by the electrolysis of molten sodium chloride to deposit Na metal at cathode by Down's process.

Many metals are purified into pure metals by the process of electrolysis

for example impure copper (blister Cu) is purified by the process of electrolysis. In this process, the impure copper i.e. Blister copper is made as anode in the electrolytic cell, cathode is a thin plate or sheet of pure copper metal and the two electrodes are dipped in the electrolytic solution of copper sulphate ($CuSO_4$), containing few drops of sulphuric acid. The two electrodes are connected with a battery (source of electricity), when an electric current is passed through the electrolytic solution. The copper anode dissolves in the solution as Cu^{2+} ions which move towards cathode and gain electrons to get neutralized, depositing pure copper metal over cathode plate. Most of the impurities of anode fall to the bottom of the cell, called as "anode mud". Copper thus deposited at cathode is 99.99% pure. In this way copper anode (Blister copper) dissolves completely to form pure copper at cathode. This process of electrolysis is similar to electroplating.

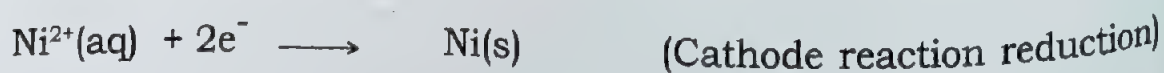
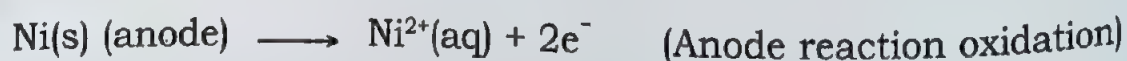
Electroplating:

Electroplating is the process of electrolysis which is used to coat one metal onto another. Usually the object to be electroplated is made up of cheaper or baser metal, such as iron, steel etc. It is then coated with a thin layer of more attractive corrosion - resistant and costly metal, like silver, gold, nickel, chromium, tin etc. The cost of the finished product is far less than the objects entirely made of these metals. Gold coated object is much cheaper than the gold object.

1. Nickel Plating:

A cell for electroplating of nickel is shown in the fig 8.4. A piece of pure nickel is the anode and the spoon or any object to be nickel plated is cathode. A solution of nickel sulphate ($NiSO_4$) is used as the electrolyte in the electrolytic cell. The two electrodes are joined with a battery (an external source of direct electric current). On passing the electric current, the anode which is Ni, dissolves in the electrolytic solution forming Ni^{2+}ions by the loss of electrons. Ni^{2+}ions from the solution move towards the cathode, where they gain electrons and are reduced to Ni metal on the surface of spoon (cathode).

$$Ni(s)\ (anode) \longrightarrow Ni^{2+}(aq) + 2e^- \quad \text{(Anode reaction oxidation)}$$

$$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s) \quad \text{(Cathode reaction reduction)}$$

The net reaction is simply the transfer of Ni as Ni^{2+} through $NiSO_4$ solution towards the cathode i.e. spoon and get it coated with Ni metal on

the surface. The sum of reduction and oxidation is :

$$Ni(s)\ (anode) \longrightarrow Ni(s)\ \ (cathode)$$

Battery

Anode (Pure Ni)

Cathode (Spoon)

Electrolyte cell

Ni^{2+} SO_4^{2-}

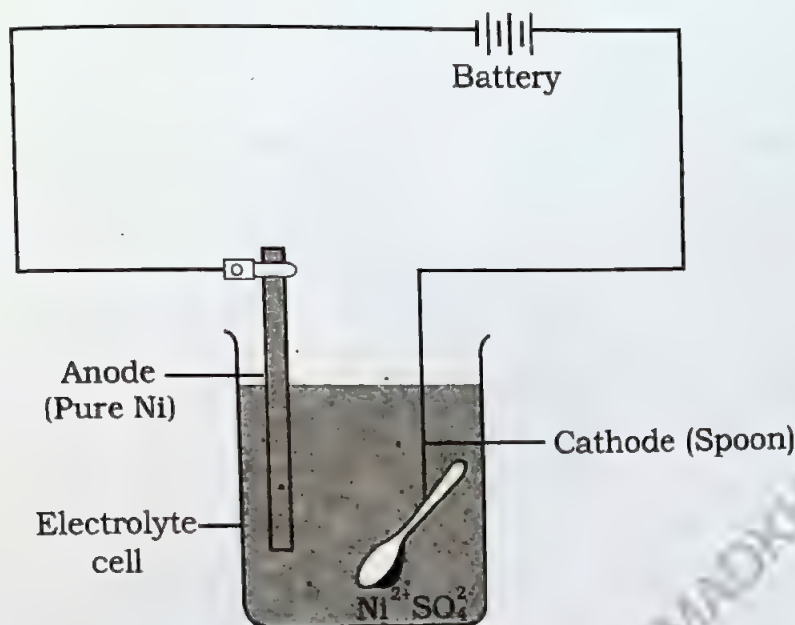
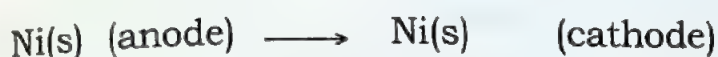



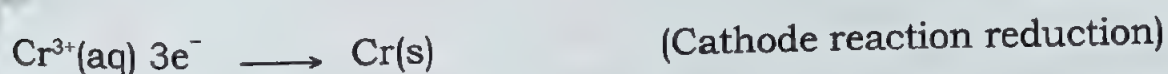
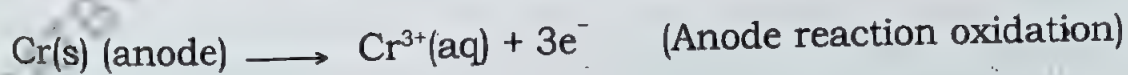
Fig. 8.4 Nickel plating

2. Chromium Plating:

Chromium metal can also be coated over baser or cheap metals by the process of electrolysis i.e. electroplating.

A cell of chromium electroplating is shown in fig. 8.5. A piece of chromium metal is the anode and the spoon or any other object to be chromium plated is the cathode. A solution of chromium sulphate [$Cr_2(SO_4)_3$] is used as an electrolyte in the electrolytic cell. The two electrodes are joined with a battery (an external source of direct electric current). On passing electric current, the anode which is chromium dissolves in the solution, forming Cr^{3+} ions by the loss of electrons, Cr^{3+} ions from the solution move towards the cathode, where they gain electrons and are reduced to deposit chromium metal on the surface of cathode.

$$Cr(s)\ (anode) \longrightarrow Cr^{3+}(aq) + 3e^- \quad \text{(Anode reaction oxidation)}$$

$$Cr^{3+}(aq)\ 3e^- \longrightarrow Cr(s) \quad \text{(Cathode reaction reduction)}$$

The net process is simply the transfer of Cr as Cr^{3+} ions through $Cr_2(SO_4)_3$ solution towards cathode and coating it with Cr-metal. The sum of oxidation and reduction is :

Cr(s) (anode) ⟶ Cr(s) (cathode)

Battery

Anode (Pure Cr)

Electrolyte cell

Cathode (Spoon)

$Cr_2(SO_4)_3$

$Cr_2(SO_4)_3$

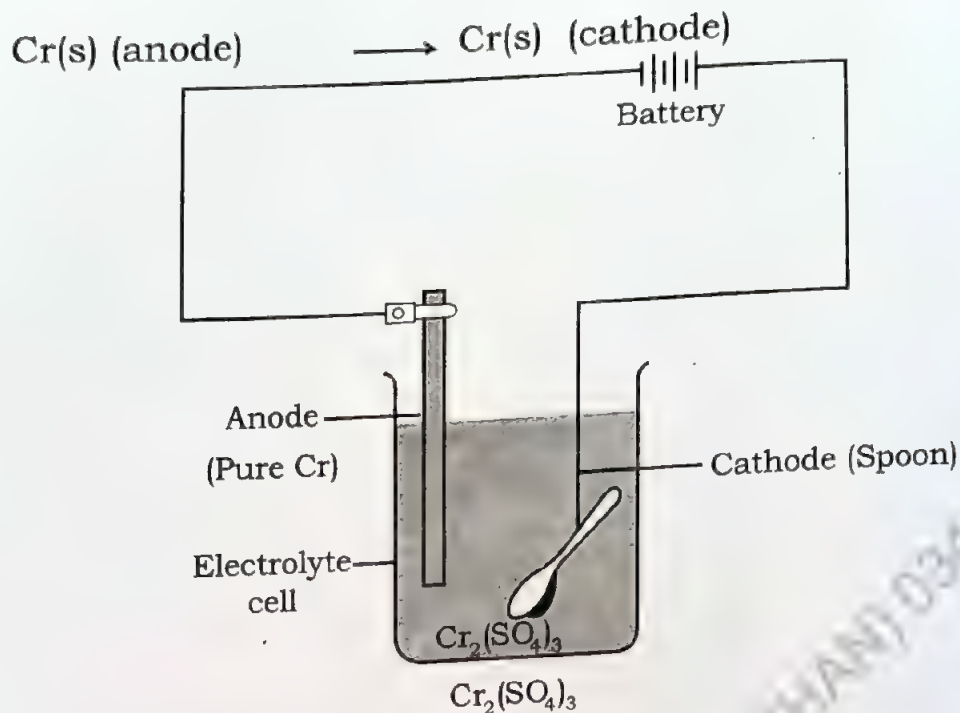



Fig. 8.5 Chromium plating

Looking for Electro plate:

Look around in your home for articles that are electroplated. Decide as far as you can :

1. What metal is plated on top.
2. What metal is underneath.
3. What is the purpose of the plating ?

Choose one of the items and look very closely at it, specially look at cooking oil container. Is there any place where the plating is noting good?

8.5 ELECTROCHEMICAL CELLS:

The cell which is used to convert chemical energy into electrical energy or vice versa is called electrochemical cell. An electrochemical cell which converts chemical energy into electrical energy is known as Galvanic or voltaic cell. This is a very strange device to produce electric current.

Take strips of zinc and copper metals, sticking them into a lemon, producing electricity. Howdoes this happen? The two dissimilar metal strips (Zn and Cu) and the electrolyte of lemon juice are the key components of the device that convert chemical energy into electrical energy. Try it on your own. Insert the Zn and Cu strips into a lemon and join the two strips by a wire outside, you will feel the electricity passing through wire joining the two strips.

The simplest of the Galvanic or Voltaic cells is Daniell cell.

Daniell Cell:

A Daniell cell is the simplest of the Galvanic or Voltaic cells which is used to convert chemical energy into electrical energy spontaneously. Daniell cell consists of two half cells. One half cell is Zinc rod (Zn - metal) dipped in IM $ZnSO_4$ solution and the other half cell is Copper rod (Cu-metal) dipped in IM $CuSO_4$ solution. The two half cells or single electrodes are connected together to form a complete cell. The two half cells should be separated from each other by a porous partition or a salt bridge, when both electrodes are connected externally through a volt-meter by means of metal wire. The cell starts producing electric current at once. Zn undergoes oxidation to form Zn^{2+} ions by the loss of 2 - electrons to go into $ZnSO_4$ solution. Zn acts as anode or negative electrode. The electrons which are free at Zn - electrode travel through the wire externally to Cu - electrode. These electrons are accepted by Cu^{2+}ions of $CuSO_4$ solution and Cu^{2+}ions undergo reduction to deposit copper metal at Cu - electrode which acts as cathode or positive electrode. In this process Zn - electrode dissolves in the solution of $ZnSO_4$ and reduces in size, while copper electrode grows in size due to the deposition of Cu - metal.

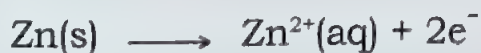
VERY SURPRIZING

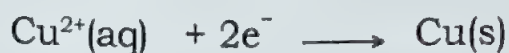
Heart Beats and Electrocardiography (ECG)

The human heart, generally in a whole day pumps more than 7000 litres of blood through the circulatory system. We generally think of the heart as a mechanical device, a muscle that circulates blood via regularly spaced muscular contractions. However, in the late 1800 two pioneers in electricity Luigi Galvani and Alesandro Volta discovered that the contractions of heart are controlled by electrical phenomenon, as are nerve impulses throughout the body. These electrical impulses that cause the contraction of heart muscle are strong enough to be detected at the surface of the body. This observation formed the basis for electrocardiography (ECG). It is quite striking that although the heart's major function is the mechanical pumping of blood, it is most easily monitored by using the electrical impulses generated by volatic cell.

Cell Reaction

At Zn - Electrode (Anode)

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

At Cu - Electrode (Cathode)

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

The total cell reaction is the sum of two half cell reactions.

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

The function of salt bridge or porous partition is to prevent the mixing of two solutions ($ZnSO_4$ and $CuSO_4$) and allows the ions to move through from one part to another. Zn^{2+} ions from the anode compartment move into the cathode compartment, while negative ions: SO_4^{2-} ions move from cathode compartment to anode compartment through the porous partition or salt bridge. It maintains the electrical neutrality in the two electrolytic solutions.

The cell voltage in Daniell cell is 1.10 volt.

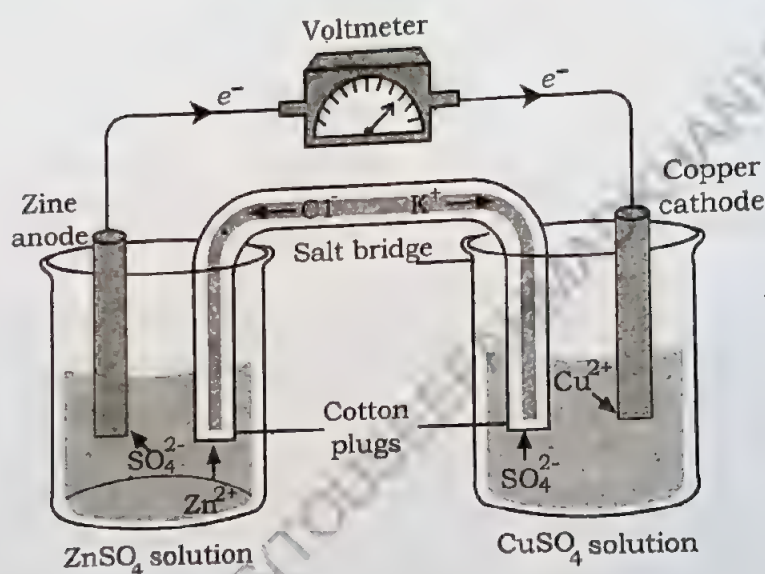



Fig.8.6 Daniell cell (A Zn - Cu voltaic cell)

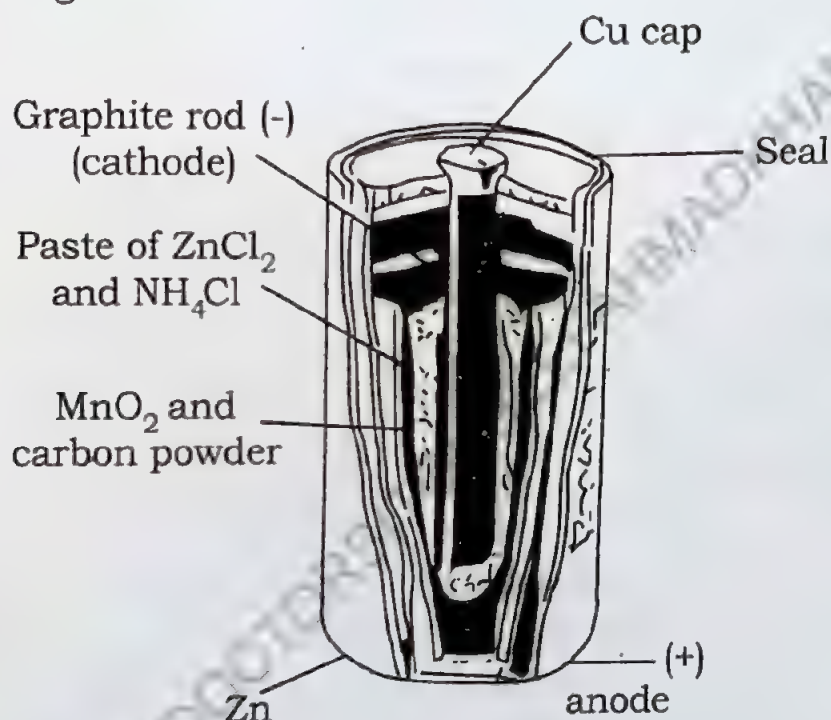
8.6 Batteries:

In every day life, we use the devices to produce electricity by the chemical reactions, known as batteries. A flash light "battery" consists of single voltaic cell with two electrodes in contact with one or more electrolytes. Some times a distinction is made between the terms "cell" and "battery". A battery is an assembly of two or more voltaic cells, connected together in series. By this definition, automobile or motor battery is a true battery. The most common types of cells or batteries are described as follows:

1. Dry Cell:

It is a primary cell, which is used to convert chemical energy into electrical energy. It is used in most of the flash lights, calculators, clocks, transistors and in portable electronic devices. It is an irreversible cell.

In a dry cell (fig 8.7), there is an outer zinc vessel which acts as anode and inert carbon (graphite) rod which acts as cathode. The graphite rod is surrounded by a mixture of manganese dioxide (MnO_2) and carbon powder. The electrolyte is a moist paste of ammonium chloride (NH_4Cl) and zinc chloride ($ZnCl_2$). The cell is called a dry cell because there is no free flowing liquid. The concentrated electrolytic solution is thickened into a gel like paste by an agent such as starch. The upper top position of the cell is sealed with wax (sealing material). A copper cap is fitted on the top of carbon rod (cathode) to make the electrical contact. The whole cell is covered with a safety cover.

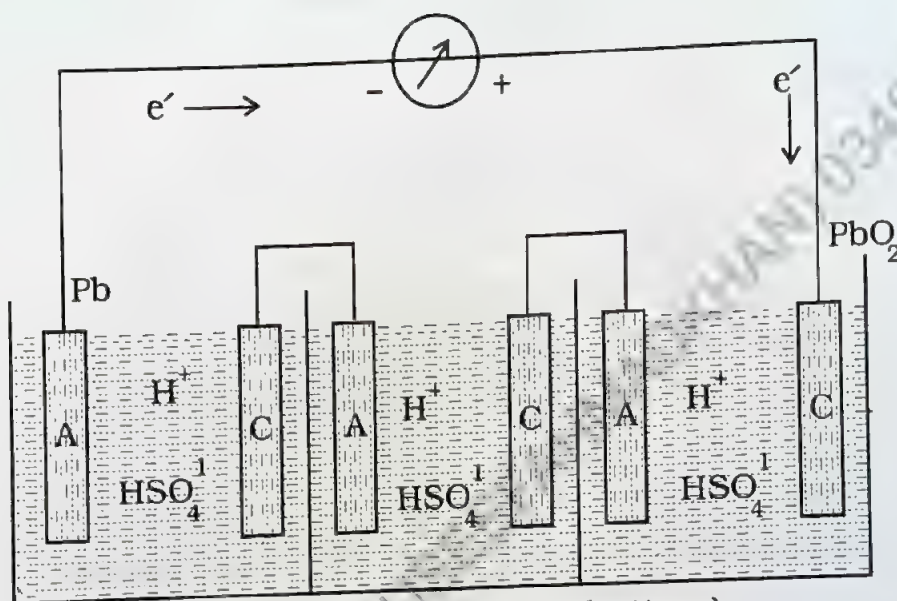
The cell diagram is as :

(Fig. 8.7 Dry cell)

When zinc and graphite electrodes are connected by a metallic wire, Zn gets oxidized to form Zn^{2+}ions which pass into the wet paste leaving behind electrons on the Zn container and the electrons move from Zn electrode to carbon electrode through the external circuit. The cell reactions are complex.

2. Lead-Storage Battery (Motor - Battery)

Lead storage battery is used in automobiles. It is a secondary battery and is a reversible cell which can be restored to its original condition. The battery can be used through repeated cycles of discharging and recharging.

Fig 8.8 shows a portion of the cell in the lead storage battery. There are several anodes and several cathodes which are connected together in series; about six cells are connected together. Each cell has a voltage of 2V and overall voltage when six cells are connected together in series would be 12V.

In lead storage battery, anodes are the lead alloy and cathodes are made up of red lead dioxide (PbO_2). The electrolyte is dilute sulphuric acid having concentration of 30%

(Fig. 8.8 Lead storage battery)

As the cell reaction proceeds, $PbSO_4$(s) precipitates and partially coats both the electrodes, and the water formed dilutes the sulphuric acid. The battery is said to be discharged. Now by connecting the battery to an external electrical source, we can force the electrons to flow in the opposite direction, i.e. the net cell reaction can be reversed and the battery is recharged.

SUMMARY

1. The chemical compounds which conduct electricity and as a result decompose by the loss and gain of electrons are called electrolytes. The chemical compounds which do not conduct electricity are called non - electrolytes.
2. Electrolytic cells are the cells, use external voltage to push a reaction non - spontaneously and are used to convert electrical energy into chemical energy while electrolysis may be defined as a process in which movements of the ions take place towards

their respective electrodes to undergo chemical changes under the influence of applied electric field. In the electrolysis of molten sodium chloride, sodium metal is formed at cathode, while chlorine gas is liberated at anode.

3. Electrochemical voltaic cell uses spontaneous chemical reactions to generate an electric current through oxidation reduction reactions. The half reactions take place in half cells. The half cell in which reduction occurs is called cathode; the half cell in which oxidation occurs is called anode. Electrons flow in the external circuit from anode to cathode. Voltaic cells are used commercially as portable energy sources (batteries). Electrolysis is an important process used for the extraction of certain metals, in purification of metals and in electroplating. Nickel - plating and chromium plating are very important techniques to protect baser metals from corrosions.

EXERCISE

1. Fill in the blanks:

(i) The substance used for electrolysis is called _________.
(ii) When molten sodium chloride is electrolysed __________ is formed at cathode.
(iii) One Faraday is equivalent to _________ coulombs.
(iv) The electrolyte in lead storage battery is _________.
(v) Dry cell is a _________ cell.

2. Write true or false:

(i) Sugar is an electrolyte.
(ii) Electrolytic conductance is also known as electrolysis.
(iii) Z is called as electrochemical equivalent.
(iv) The unit of electrochemical equivalent is ampere x second.
(v) Daniell cell is a voltaic cell.
(vi) Lead storage battery is a primary cell.

3. Write answers of the following questions:

(i) Define the following terms:

(a) Electrolysis
(b) Electrochemical cell
(c) Coulomb
(d) Electrochemical equivalent
(e) Primary cell
(f) Electroplating

(ii) Predict what would be formed (i) at the anode and (ii) at the cathode when each of the following molten salts are electrolysed using inert electrodes.

(a) $NaCl$ (b) $MgBr_2$ (c) $CaCl_2$

(iii) State and explain Faradays first law of electrolysis.

(iv) Calculate the amount of silver deposited when 10 ampere of current is passed for 50 minutes through a solution of $AgNO_3$. (Z of Ag = 0.00118 g/C)

(v) Describe the construction and working of a dry cell.

(vi) What happens when electric current is passed through acidulated water? Give the reactions at two electrodes and mention the products at cathode and anode.

(vii) Predict the net electrolysis reaction when molten NaCl is electrolysed.

(viii) Describe the process of nickel plating.

(ix) What is the function of a salt bridge or porous partition in an electrochemical cell?

(x) Which of the following pairs of terms have the same meaning and which have a different meaning ?
   (a) Voltaic cell and Galvanic cell.
   (b) Electrolytic cell and Electrochemical cell.
   (c) Cell and battery.

(xi) What is the difference between a primary and a secondary cell? Give an example of each. Discuss lead - storage battery.

(xii) When molten NaCl is electrolysed. Sodium metal is liberated at cathode by the reaction $Na^+ + e^- \longrightarrow Na(s)$. How many grams of sodium are liberated when 5×10^3 C of electric charge is passed through the cell.

CHAPTER-9

ACIDS, BASES AND SALTS

You will learn in this chapter about:

- Acids and bases.
- Arrhenius theory.
- Bronsted lowry theory.
- Lewis concept.
- Properties of acids and bases.
- Neutralization.
- Basicity and acidity.
- Strength of acids and bases.
- Salts.
- Classification of salts and some important salts.
- Dissociation of water.
- The concept of pH and pH scale.
- The importance of pH.
- Acid-base titration.
- Standard solution.

9.1 ACIDS AND BASES:

Introduction:

By the 15th century, chemists recognized a group of substances which had sour taste called acids (In Latin acidus meaning 'sour'). They also recognized another group of substances which had bitter taste and

were used as good cleaning agents, called Bases.

In the 16th century, it was recognized that bases react with acids to 'destroy' or "neutrailize" them, forming an ionic compound called salt.

Nearly all fruits and foods contain acids. Lemons, oranges, grape-fruits, contain citric acid. All citrous fruits contain large amount of ascorbic acid ($C_6H_8O_6$), or vitamin-C. Ascorbic acid also acts as an anti-oxidant. Apples contain maleic acid. The souring of milk produces lactic acid, butter on rancidity gives butyric acid. The extract of vinegar is acetic acid. Chemists prepare large quantities of important industrial acids. They are manufactured from minerals and are known as inorganic acids or more commonly mineral acids. These include hydrochloric acid (HCl), nitric acid (HNO_3), sulphuric acid (H_2SO_4) and phosphoric acid (H_3PO_4). The most important acid is H_2SO_4. The consumption of (H_2SO_4) is an index to the state of civilization and prosperity of a country. The important acid for making explosives and fertilizers is (HNO_3), and (HCl) is used as cleaning agent. It composes about 0.4% of gastric juice of our stomach and aids in digestion of food.

There are several substances found in almost every home, called bases. House hold ammonia (NH_3+H_2O solution) is a common cleaning agent. Lye is commercial (NaOH) used for cleaning, sink–drains. Lime water is a solution of $Ca(OH)_2$. Milk of magnesia ($Mg(OH)_2$) is used as an antacid, laxative, and an antidote, when strong acids are swallowed.

Salts have a positive ion other than (H^+) combined with a negative ion, other than (OH^-). However, one must keep in mind that all the salts are not neutral, some behave like acids and others like bases. Thus many substance were grouped into one of the three classes, acids, bases, and salts.

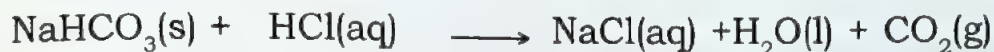
9.2 PROPERTIES OF ACIDS AND BASES

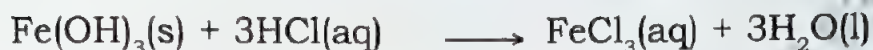
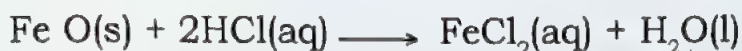
Acids:

1. Acids have a sour taste, vinegar gives the taste of acetic acid and lemons and other citrous fruits contain citric acid.
2. They change the colour of blue litmus to red.
3. Acids react with certain metals, such as (Zn, Mg and Fe) to produce (H_2) gas, for example, when diluted, (HCl) reacts with (Zn) metal, producing (H_2) gas.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(s) + H_2(g)$$

4. Acids react with carbonates and bicarbonates, such as Na_2CO_3, $CaCO_3$ and $NaHCO_3$ to produce CO_2 gas.

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$
$$NaHCO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

5. Acids react with oxides and hydroxide of metals, forming salt and water.

$$FeO(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2O(l)$$
$$Fe(OH)_3(s) + 3HCl(aq) \longrightarrow FeCl_3(aq) + 3H_2O(l)$$

6. Aqueous acid solutions, conduct electricity.
7. They react with bases to form salts and water.

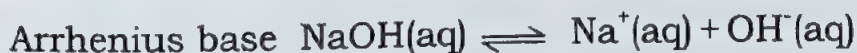
Bases:

1. Bases have bitter taste.
2. Bases have slippery touch.
3. They change the colour of red litmus to blue
4. Aqueous solutions of bases, conduct electricity.
5. They react with acids to form salts. When acids and bases are mixed in right proportions, the acidic and basic properties disappear and new substances salt and water are formed and the reaction is neutralization reaction.

9.2.1 Acids and Bases (Arrhenius theory):

Svante-Arrhenius, a Swedish chemist in (1887) first gave the clue to chemical nature of acids in his theory of ionization. It states, that an acid can be defined as a substance, that yields hydrogen (H^+) ions when dissolved in water. A base can be defined as a substance that yields hydroxide (OH^-) ions, when dissolved in water. Thus (HCl) and (H_2SO_4) are acids and (NaOH) and (KOH) are bases.

Arrhenius acid $HCl(aq) \rightleftharpoons H^+(aq) + Cl^-(aq)$

Arrhenius base $NaOH(aq) \rightleftharpoons Na^+(aq) + OH^-(aq)$

Bronsted Lowry Theory (Proton-Donor and Acceptor Theory):

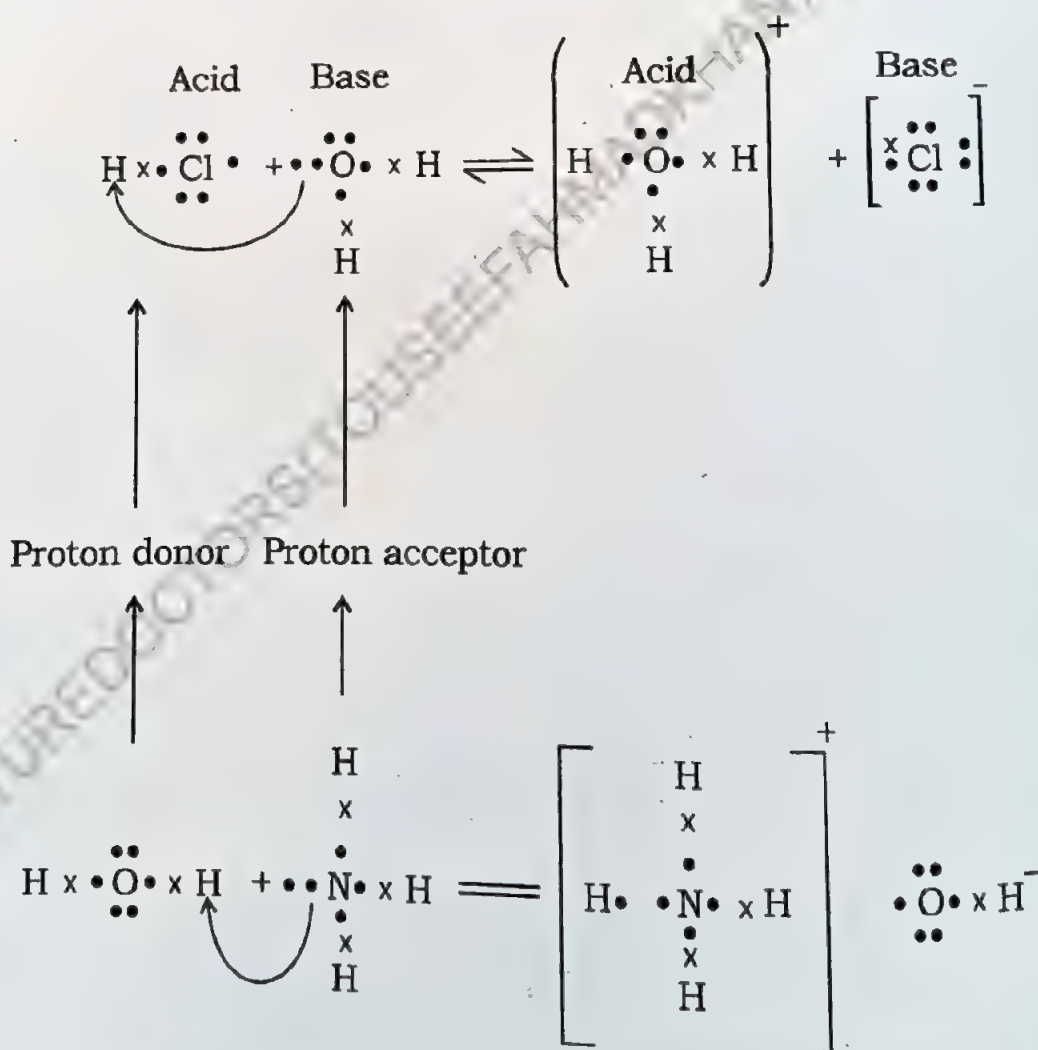
Arrhenius's definitions of acids and bases are limited in that, they apply only to water (aqueous) solutions and it also does not account for the basicity of ammonia (NH_3), that doesn't contain (OH) group. Broader

definitions, which were proposed by the Danish Chemist Johannes Bronsted and English Chemist Thomas Lowry, in 1923, describe, that *An acid is a substance having a tendency to donate one or more protons and base is a substance, having a tendency to accept (add) protons.*

Bronsted-Lowry acid = A substance that can donate (H^+)
Bronsted-Lowry base = A substance that can accept (H^+)

In the above examples, the (HCl) and water (H_2O), are proton donors and act as Bronsted-Lowry acids, whereas (H_2O) and ammonia (NH_3) are proton acceptors and are known as Bronsted Lowry bases.

Acid Base Acid Base

Proton donor Proton acceptor

LEWIS CONCEPT OF ACIDS AND BASES

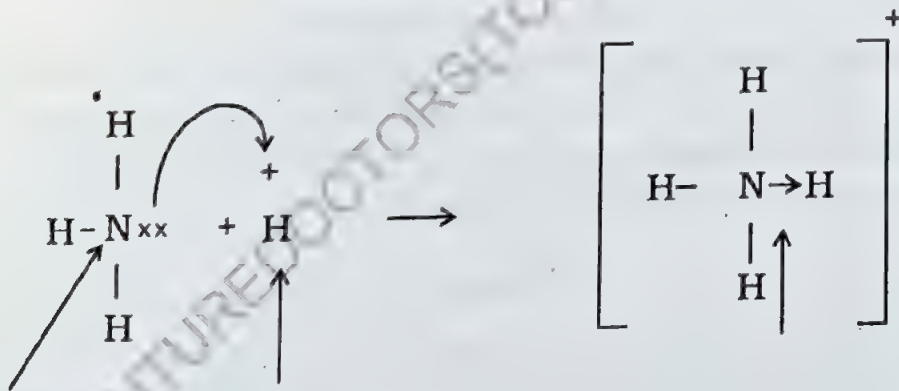
(Electronic explanation of Acids and Bases):

In 1923 G. N. Lewis proposed a more general concept of acids and bases. It explains the behaviour of acids in terms of electrons rather than protons, since electrons are responsible for chemical bonding. Lewis acids include not only (H^+) protons, but also other cations and neutral molecules, having vacant valence orbitals that can accept a pair of electrons donated by Lewis base.

According to the Lewis theory, an acid is any species (molecule or ion) which can accept a pair of electrons, and base is any species (molecule or ion) which can donate a pair of electrons. An acid-base reaction, in which electron pair donor is base and electron-pair acceptor is acid, they form a co-ordinate covalent bond between the two.

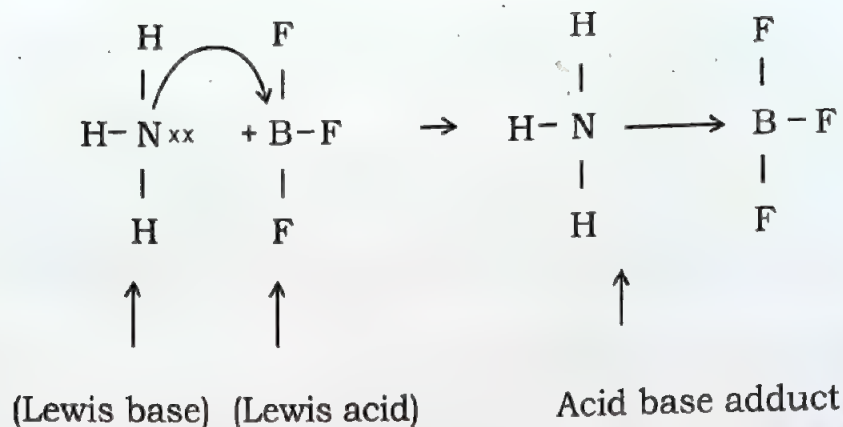
Lewis Acid = An electron pair acceptor.
Lewis Base = An electron pair donor.

For Example: When ammonia (NH_3) reacts with proton (H^+) to form ammonium ion (NH^+_4), in which the nitrogen of (NH_3) donates a pair of electrons whereas the (H^+) accepts that pair of electrons for bond formation, this is shown by curved arrow.

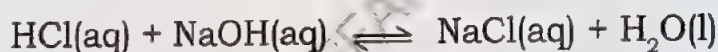
$$H_3N^{\times\times} + H^+ \longrightarrow \left[H_3N \rightarrow H \right]^+$$

Electron pair donor (Lewis base) — Electron pair acceptor (Lewis acid) — Co-ordinate covalent bond

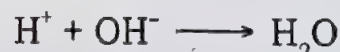
Another example is provided by the reaction of ammonia (NH_3) with boron tri-fluoride (BF_3), in which nitrogen of (NH_3) donate an electron pair and (B) of BF_3, which lacks a pair of electrons to complete its outer most shell (octet), accepts that pair of electrons and form a co-ordinate covalent bond.

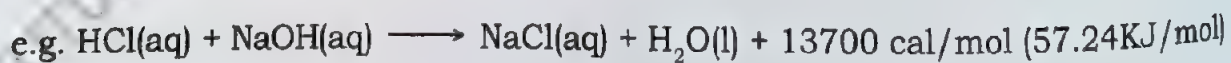
$$H_3N: + BF_3 \rightarrow H_3N \longrightarrow BF_3$$

(Lewis base) (Lewis acid) Acid base adduct

9.2.2 Neutralization

If we add an acid to a base drop by drop, the acidic character of the acid decreases gradually. A stage will come when the resultant solution will become neutral to litmus. This stage is called neutralization. In other words when equivalent quantities of an acid and a base are mixed, salt and water are formed. A common example is the reaction between HCl and NaOH.

$$HCl(aq) + NaOH(aq) \rightleftharpoons NaCl(aq) + H_2O(l)$$

The hydrogen ion (H^+) which is responsible for acidic properties, reacts with the hydroxide ion (OH^-), which is responsible for the basic properties, producing neutral water (H-OH). Because the only change that takes place is the reaction of the hydrogen (H^+) ion and hydroxide ion (OH^-), the neutralization may be expressed simply as:

$$H^+ + OH^- \longrightarrow H_2O$$

The neutralization is an exothermic reaction.

e.g. $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l) + 13700$ cal/mol (57.24KJ/mol)

In case either acid or base is not completely ionized, i.e. (weak acid or weak base), the neutralization reaction may not go to completion, and the heat of neutralization may be less than 13700 cal. For example, when strong acid (HCl) reacts with weak base Ca $(OH)_2$, the amount of heat evolved is 24700 cal/mol.

$$2\,HCl + Ca\,(OH)_2 \longrightarrow CaCl_2 + 2H_2O + 24700 \text{ cals}$$

$H^+ + Cl^-$

$Na^+ + OH^-$

$Na^+ + Cl^- + HOH$

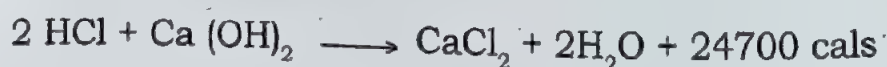
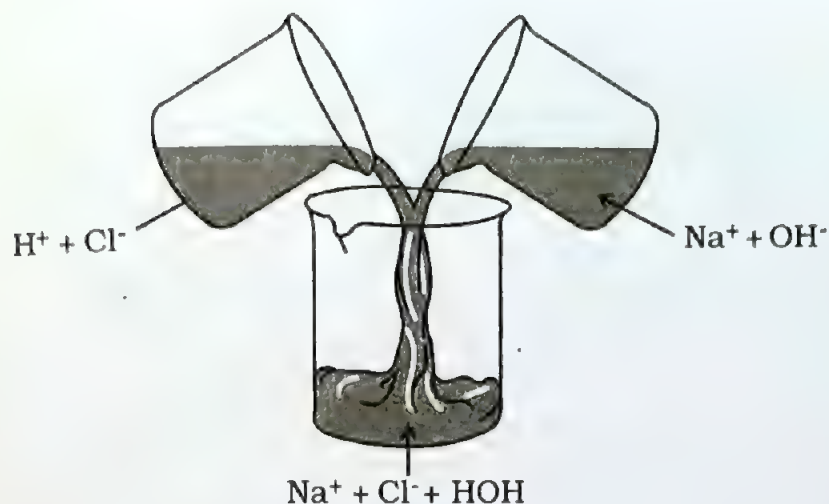



Fig. 9.1 Shows when solutions of hydrochloric acid and sodium hydroxide are mixed, only hydrogen ions and hydroxide ions react forming water. Therefore, we may write the equation: $H^+(aq) + OH^-(aq) \rightarrow HOH(l)$

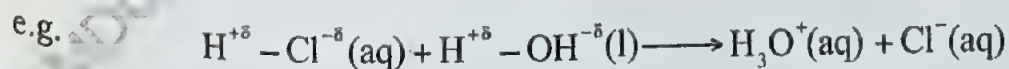
REMEMBER

A reaction in which an acid and a base form an ionic compound (salt and water) is called neurtralization reaction. If water is formed, the reaction can be classified as a double-replacement reaction.

9.2.3 Mono and Poly Acids and Bases (Mono Basic and Poly Basic Acids) (Basicity of Acids)

As we have seen that an acid yields the protons. Different acids have different number of protons (acidic-hydrogen) per molecule and yield different number of (H_3O^+) ion in a solution.

The common acids like (HCl) (HNO_3) and (CH_3COOH) contain only one acidic hydrogen atom per molecule. When dissolved in water, 1 mole of each of these acids is capable of producing 1 mole of hydrated H^+ = (H_3O^+) ions, and in order to neutralize this solution,1 mole of (OH^-)ions is required. Consequently these acids are called mono-basic acids, more commonly called mono-protic acids

e.g. $$H^{+\delta}-Cl^{-\delta}(aq)+H^{+\delta}-OH^{-\delta}(l)\longrightarrow H_3O^+(aq)+Cl^-(aq)$$

The number of replaceable or ionizable hydrogen atoms present in a molecule of an acid is called 'Basicity" of the acid.

Sulphuric acid (H_2SO_4) contains two acidic hydrogen atoms per molecule. It can neutralize two molecules of hydroxide (OH^-) ions. Consequently sulphuric acid (H_2SO_4) is called dibasic acid (Diprotic acid). Sulphuric acid (H_2SO_4) dissociates in two steps :

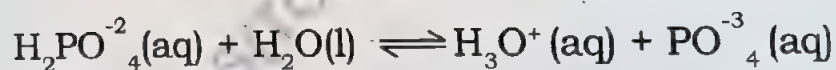
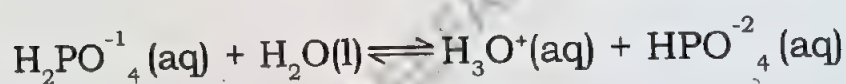
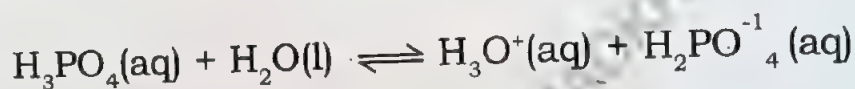
$$H_2SO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$$

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{-2}(aq)$$

The acid like phosphoric acid (H_3PO_4), which contains three acidic hydrogen atoms per molecule can neutralize three molecules of (OH^-)ions, so phosphoric acid (H_3PO_4) is called tri basic acid or (Tri-protic acid).

$$H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + H_2PO_4^{-1}(aq)$$

$$H_2PO_4^{-1}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{-2}(aq)$$

$$H_2PO_4^{-2}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + PO_4^{-3}(aq)$$

Acids that contain two, three or more acidic hydrogen per molecule, are called poly-basic acids, or more commonly poly-protic acids.

9.2.4 Mono-Acid and Poly-Acid Bases (Acidity of Bases):

Similarly, bases that produce 1 mole of (OH^-) ions per mole of base (such as NaOH and KOH) are called mono-acid bases. Bases that produce 2 moles of (OH^-) ions per one mole of base (such as $Ca(OH)_2$ and $Ba(OH)_2$) are called di-acid bases, and bases that produce 3 moles of (OH^-) ions per 1 mole of base (such as $Al(OH)_3$ and $Cr(OH)_3$ are called tri-acid bases.

The number of ionizable or replaceable ($-OH^-$) ions, present in a molecule of base is called acidity of the base.

Bases that contain two three or more hydroxide (OH^-) ions per molecule are called poly-acid bases,

9.3 DISSOCIATION OF ACIDS AND BASES (Acid Strength and Base Strength)

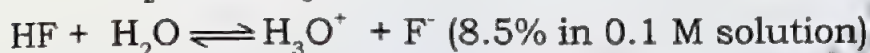
Different acids differ in their ability to donate protons. A strong acid is one, that is almost completely dissociated (strong electrolytes) i.e. an acid that produces large number of (H^+) ions in aqueous solution is said to be a strong acid. Typical examples of strong acids are hydrochloric acid (HCl) nitric acid (HNO_3) and sulphuric acid (H_2SO_4).

A weak acid is one, that is only partially dissociated (weak-electrolytes). Only a small fraction of the weak acids transfer a proton to water. Typical examples of weak acids are nitrous acid (HNO_2), phosphoric acid (H_3PO_4), hydro fluoric acid (HF) carbonic acid (H_2CO_3) acetic acid (CH_3-COOH) and formic acid (HCOOH).

For example: In strong acid, greater is the extent of ionization in water.

$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$ (complete ionization in dilute solution)

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$ (8.5% in 0.1 M solution)

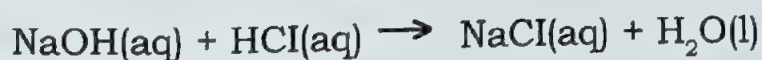
$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$ (1.3% in 0.1 M solution)

Similarly, the strong base is one, that is almost dissociated completely (strong electrolytes), that is, a base which yields large number of (OH^-) ions in aqueous solution. Most metal hydroxides, such as NaOH, KOH, $Ba(OH)_2$ and $Ca(OH)_2$ are strong electrolytes and strong bases.

A weak base is one, that is only partially dissociated (weak electrolytes). Weak bases dissociate to a small extent in water to yield (OH^-) ions.

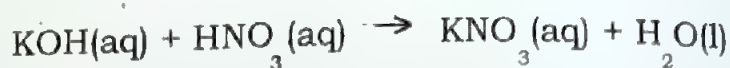
Typical examples of weak bases are, NH_4OH, $Mg(OH)_2$ and $Be(OH)_2$, etc. Hence, the relative strength of weak bases may be measured by the extent to which they dissociate in water to yield hydroxide ions (OH^-).

9.4 SALTS

A salt is ionic compound produced when an acid is neutralized by a base. For example, sodium hydroxide neutralizes hydrochloric acid to form sodium chloride (salt) and water.

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Similarly potassium hydroxide neutralizes nitric acid to form

potassium nitrate and water.

$$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$$

On the basis of their chemical nature, salts can be divided into three groups.

1. Normal Salts 2. Acidic salts. 3. Basic salts.

1. Normal Salts:

Salts which are formed by the complete neutralization of an acid by a base e.g. NaCI, $NaNO_3$, K_2SO_4 etc are normal salts. These salts do not have replaceable hydrogen atoms or hydroxyl groups.

2. Acidic Salts:

Salts which are formed by the partial neutralization of an acid by a base e.g. $NaHSO_4$, $KHCO_3$, etc are acidic salts.

$$NaOH(aq) + H_2SO_4(aq) \longrightarrow NaHSO_4(aq) + H_2O(l)$$

$$KOH(aq) + H_2CO_3(aq) \longrightarrow KHCO_3(aq) + H_2O(l)$$

These salts contain replaceable hydrogen ion. They react further with bases to form normal salts.

3. Basic Salts:

Salts, which are formed by the partial neutralization of a base by an acid. e.g. Mg (OH) CI, Zn(OH)CI, etc are basic salts.

$$Mg(OH)_2 + HCI(aq) \longrightarrow Mg(OH)Cl + H_2O(l)$$

These salts have replaceable hydroxyl groups. They can further react with acids to form normal salts.

DOUBLE SALTS

The crystalline compounds which are obtained, when two specific salts are crystallized together are known as double salts. These salts have definite chemical composition. These compounds usually have definite number of water molecules with them. Typical examples of double salts are:

Potash Alum	$K_2SO_4 . Al_2 (SO_4)_3 . 24H_2O$
Chrome Alum	$K_2 SO_4 . Cr_2(SO_4)_3 . 24H_2O$
Carnalite	$KCl. MgCl_2 . 6H_2O$
Mohr's Salt	$FeSO_4 . (NH_4)_2 SO_4 . 6H_2O$

9.4.1 Some important Commercial Preparation and uses of Salts.

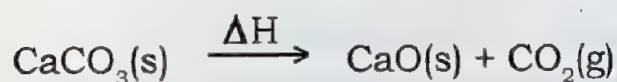
1. Sodium carbonate: ($Na_2CO_3.10 H_2O$)

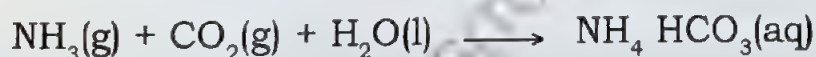
Today sodium carbonate (Na_2CO_3) is commercially prepared by the **Solvay process** or ammonia soda process.

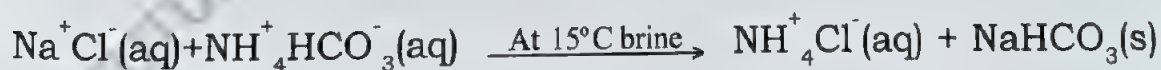
The raw materials are lime stone ($CaCO_3$), sodium chloride (NaCl), ammonia (NH_3) and water.

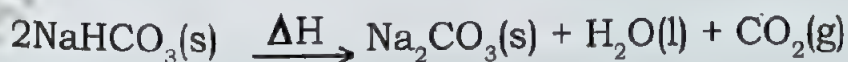
The process involves the following steps:

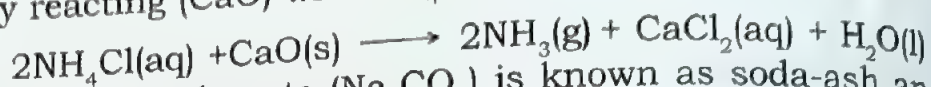
[i] Lime stone [$CaCO_3$] is heated to yield calcium oxide (quicklime CaO) and the CO_2 gas.

$$CaCO_3(s) \xrightarrow{\Delta H} CaO(s) + CO_2(g)$$

[ii] This (CO_2) is passed into aqueous solution of ammonia, and the ammonium bicarbonate is produced.

$$NH_3(g) + CO_2(g) + H_2O(l) \longrightarrow NH_4 HCO_3(aq)$$

[iii] This (NH_4HCO_3) reacts with aqueous cold solution of (NaCl) at 15°C, called **Brine** to yield, sodium bicarbonate ($NaHCO_3$), which is not soluble at low temperature (15°C) and this precipitates out.

$$Na^+Cl^-(aq)+NH^+_4HCO^-_3(aq) \xrightarrow{\text{At } 15^\circ C \text{ brine}} NH^+_4Cl^-(aq) + NaHCO_3(s)$$

[iv] This ($NaHCO_3$) on heating yields sodium carbonate.

$$2NaHCO_3(s) \xrightarrow{\Delta H} Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

The ammonia (NH_3) which is used as a raw material in 2nd step, is recovered by reacting (CaO) with NH_4Cl.

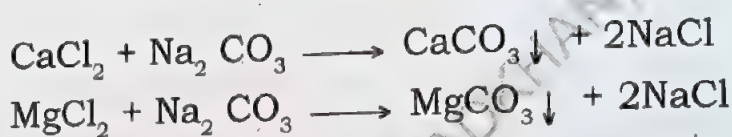
$$2NH_4Cl(aq) + CaO(s) \longrightarrow 2NH_3(g) + CaCl_2(aq) + H_2O(l)$$

Anhydrous sodium carbonate (Na_2CO_3) is known as soda-ash and sodium carbonate decahydrate ($Na_2CO_3.10H_2O$) is commonly known as washing soda.

Uses of Sodium Carbonate (Na_2CO_3):

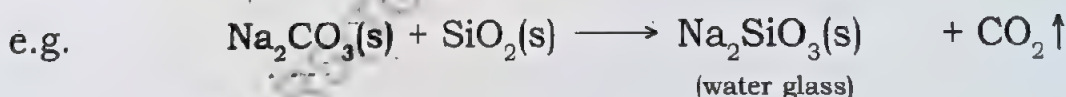
Sodium carbonate is soluble in water and has many important applications.

[a] It is used in the softening of water. Sodium carbonate furnishes carbonate ion (CO_3^{-2}) to precipitate calcium and magnesium ions,

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3\downarrow + 2NaCl$$

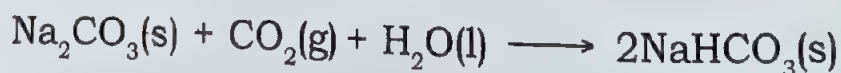
$$MgCl_2 + Na_2CO_3 \longrightarrow MgCO_3\downarrow + 2NaCl$$

[b] It is used as cleaning agent, and in making of soap, detergents and paper.

[c] It is used in making ordinary glass, which is used in bottles.

e.g. $$Na_2CO_3(s) + SiO_2(s) \longrightarrow \underset{\text{(water glass)}}{Na_2SiO_3(s)} + CO_2\uparrow$$

2. Sodium Hydrogen carbonate ($NaHCO_3$) (Baking Soda):

Preparation:

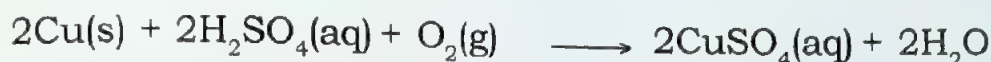
Sodium hydrogen carbonate ($NaHCO_3$) or baking soda is formed by *"Solvay process"*, but mostly it is prepared by passing the stream of CO_2 through concentrated aqueous (Na_2CO_3) solution.

$$Na_2CO_3(s) + CO_2(g) + H_2O(l) \longrightarrow 2NaHCO_3(s)$$

Uses:

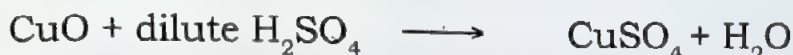
1. Baking soda is used in the preparation of baking powder
2. In the preparation of effervescent drinks and fruit salts
3. In medicines to remove acidity of stomach (i.e. as Antacid).
4. In fire extinguishers.

3. Copper Sulphate ($CuSO_4.5H_2O$):

Copper sulphate or cupric sulphate which is also known as blue vitrol or blue stone may be prepared by reacting copper scraps with dilute sulphuric acid in the presence of air.

$$2Cu(s) + 2H_2SO_4(aq) + O_2(g) \longrightarrow 2CuSO_4(aq) + 2H_2O$$

It can also be prepared by the treatment of CuO or $CuCO_3$ with dilute sulphuric acid (H_2SO_4).

$$CuO + \text{dilute } H_2SO_4 \longrightarrow CuSO_4 + H_2O$$

$$CuCO_3 + \text{dilute } H_2SO_4 \longrightarrow CuSO_4 + CO_2 + H_2O$$

Uses of Copper Sulphate:

1. In textiles (mordant), tanning, electric batteries, hair dyes and in electroplating.
2. As germicide, insecticide, preservative for wood and paper pulp.
3. In calico printing, making synthetic rubber and copper salts e.g scheels, green paint,
4. In paint and varnish industry.
5. A mixture of copper sulphate and milk lime is used to kill fungus and molds.

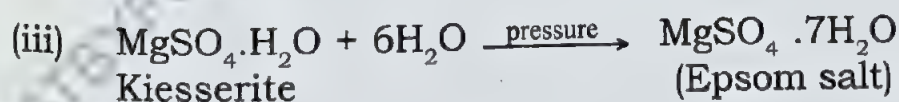
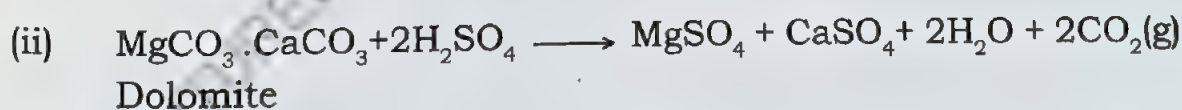
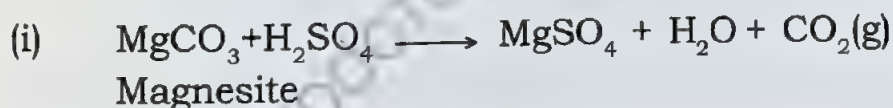
4. Magnesium Sulphate ($MgSO_4.7H_2O$) (Epsom-Salt):

Preparation:

It is prepared by the action of H_2SO_4 and magnesite or dolomite, but nowadays it is prepared by heating kiesserite under pressure with water.

(i) $\underset{\text{Magnesite}}{MgCO_3} + H_2SO_4 \longrightarrow MgSO_4 + H_2O + CO_2(g)$

(ii) $\underset{\text{Dolomite}}{MgCO_3.CaCO_3} + 2H_2SO_4 \longrightarrow MgSO_4 + CaSO_4 + 2H_2O + 2CO_2(g)$

(iii) $\underset{\text{Kiesserite}}{MgSO_4.H_2O} + 6H_2O \xrightarrow{\text{pressure}} \underset{\text{(Epsom salt)}}{MgSO_4.7H_2O}$

Uses:

1. It is used as a mild purgative in medicines.
2. In dyeing and tanning processes.
3. In making fire proof fabrics.
4. As a filler in paper industry.
5. In manufacture of ceramics, glazed tiles and match boxes.

5. **Potash Alum** **($K_2SO_4.Al_2(SO_4)_3.24H_2O$):**

If equi-molecular quantities of potassium sulphate and aluminium sulphate are dissolved in water and the solution is allowed to evaporate. crystals of ($K_2SO_4.Al_2(SO_4)_3.24H_2O$) which is called ordinary alum or potash alum are seperated out.

Uses:

1. Alum is used in dyeing as mordant to fix insoluble dye to fibre.
2. It is also used in tanning leather.
3. In sizing paper.
4. In purifying water.
5. As an antiseptic and as a mouth wash.
6. It also used in medicines.

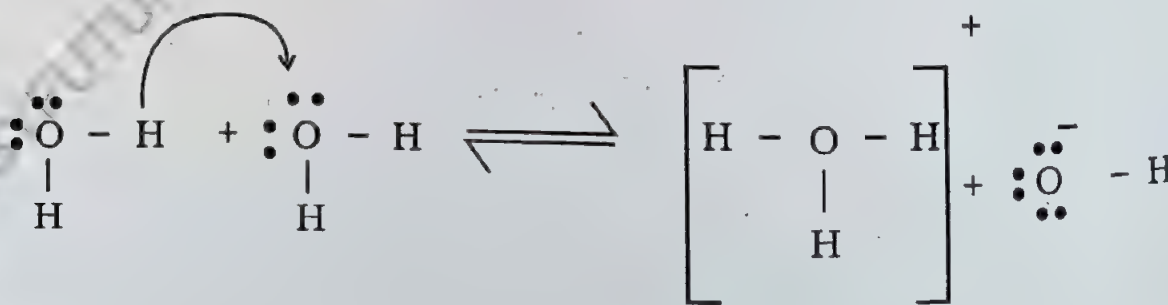
9.5 DISSOCIATION OF WATER (The acid-base properties of water):

One of the most important properties of water is its ability to act both as an acid and as a base. In presence of an acid, water acts as a base where as in the presence of base, water acts as an acid.

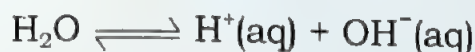
$$\text{Acid} + \text{Base} \rightleftharpoons \text{Acid} + \text{Base}$$
$$HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$$

$$H_2O(l) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

A substance (such as water) that can behave as both an acid and a base is said to be an amphoteric substance. The amphoteric nature of water is best seen in its self ionization. A proton from one water molecule is transferred to another water molecule, leaving behind (OH^-) ion and forming (H_3O^+) ion.

$$H{-}\ddot{\underset{\cdot\cdot}{O}}{-}H + H{-}\ddot{\underset{\cdot\cdot}{O}}{-}H \rightleftharpoons \left[H{-}\underset{|}{O}{-}H \atop \;\;\; H\right]^+ + :\ddot{\underset{\cdot\cdot}{O}}^-{-}H$$

9.5.1 The Ion Product of Water:

In the study of acid-base reaction in aqueous solution, the important quantity is the (H^+) ions concentration, expressing the proton as (H^+) rather than (H_3O^+). We can write the equilibrium constant (K_c) for the auto ionization water.

$$H_2O \rightleftharpoons H^+(aq) + OH^-(aq)$$

$$Kc = \frac{[H^+][OH^-]}{H_2O}$$

Since a very small fraction of water molecules are ionized, the concentration of water, that is $[H_2O]$ remains mostly unchanged, therefore,

$$K_c [H_2O] = K_w = [\overset{+}{H}] [O\overline{H}]$$

The equation constant (K_w) is called the ion-product constant, which is the product of molar concentration of (H^+) and (OH^-) ions at a particular temperature.

In pure water at 25^0C,the concentration of H^+ and OH^- ions is equal and found to be:

$[H^+] = 1x10^{-7}$ M and $[O\overline{H}] = 1x10^{-7}$ M

Thus $K_w = (1x10^{-7}) \times (1x10^{-7}) = 1x10^{-14}$

Whenever $[\overset{+}{H}] = [O\overline{H}]$, the aqueous solution is said to be neutral, if the number of (H^+)ions increase, the aqueous solution is acidic. When the number of (OH^-) ions increase, the aqueous solution is basic.

For example, the (H^+) ion concentration of any solution is $1x10^{-4}$, then the (OH^-) ion concentration must change to,

$$K_w = [\overset{+}{H}] [O\overline{H}] = 1x10^{-14}$$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1\times10^{-14}}{1\times10^{-4}}$$

$$= 1\times10^{-14}\times10^{4}$$

$$= 1\times10^{-10} M$$

∴ OH^- ion concentration = 1×10^{-10} M

9.5.2 The Concept of (pH):

It is well known fact that the strength of acidic solution is measured in moles per litre of a solution. A one molar (1M) solution of HCl contains 1mole of (H_3O^+) ions in each litre (dm^3) of solution. In pure water (or any neutral solution) there are $1x10^{-7}$M (moles per litre) of H_3O^+ ions and equal number i.e. ($1x10^{-7}$M) of OH^- ions. To avoid the use of complex numbers, such as $1x10^{-7}$M or $1x10^{-10}$ M, to express the concentration of H^+ and OH^- ions, it is convenient to express the acidity or basicity of a solution in terms of pH. In 1909 the Danish Chemist S.P.L. Sorensen proposed that only the number in the exponent be used to express the acidity, called (pH) from the French (Pouvoir hydrogen = hydrogen power). On this scale, a concentration of ($1x10^{-7}$) moles of H_3O^+ ions per litre of solution becomes a pH of 7. Similarly a concentration $1x10^{-10}$M, becomes a pH of 10, and so on. *Thus (pH) of a solution is defined as the negative logarithm of the hydrogen ion (H^+) concentration or (H_3O^+) ion concentration (in moles per litre).*

Mathematically, we can write pH = -log [H^+]

Thus, a pure water (or any neutral solution) in which (H^+) ion concentration is $1x10^{-7}$ M, has a pH of 7.

$$\text{Since } pH = -\log[H^+]$$
$$pH = -\log[10^{-7}]$$
$$pH = -[-7]$$
$$\therefore \quad pH = 7$$

Similarly pOH is the negative logarithm of hydroxide ion (OH^-) concentration.

Mathematically, we can write pOH = -log [OH^-]

The sum of pH and pOH of a solution is always equal to 14.

i.e $\boxed{pH + pOH = 14}$

Hydrogen and Hydroxide Ion Concentration in Various Acidic and Basic Water Solutions at 25°C

Degree of Acidity	$[H^+]$	$[OH^-]$	Preparation	pH	pOH
Very acidic	10^{+1}	10^{+15}		-1	15
	10^{0}	10^{-14}		0	14
	10^{-1}	10^{-13}		1	13
Medium Acidity	10^{-2}	10^{-12}	Add sufficient acid to water	2	12
	10^{-3}	10^{-11}		3	11
	10^{-4}	10^{-10}		4	10
Slightly Acidic	10^{-5}	10^{-9}		5	9
	10^{-6}	10^{-8}		6	8
Neutral	10^{-7}	10^{-7}	Pure water or solutions of Neutral substance	7	7
Slightly basic	10^{-8}	10^{-6}		8	6
	10^{-9}	10^{-5}		9	5
Medium alkalinity	10^{-10}	10^{-4}		10	4
	10^{-11}	10^{-3}		11	3
	10^{-12}	10^{-2}		12	2
Very basic	10^{-13}	10^{-1}	Add sufficient base to water	13	1
	10^{-14}	10^{0}		14	0
	10^{-15}	10^{+1}		15	-1

The following examples illustrate the calculation of pH, pOH and (H^+) ion concentrations.

Examples. (1) Calculate the pH of 0.01 M HCl solution?

Solution:

We know that, HCl is strong acid, so it is completely ionized in solution.

Then $pH = -\log [H^+]$

Express $[H^+]$ in exponential form

$0.01\ M \longrightarrow 1\times10^{-2}\ M$

$pH = -\log [H^+]$

$pH = -\log [10^{-2}]$

$= -\log [-2]$

$pH = 2$ (Answer)

Example. (2) Calculate the pH of a solution whose (H^+)ion concentration is $5\times10^{-4}M$?

Solution:

$$pH = -\log [H^+]$$
$$= -(\log 5 + \log 10^{-4})$$
$$= -[.699 - 4]$$
$$= -.699 + 4$$
$$pH = 3.301 \text{ (Answer)}$$

Example. (3) Calculate the pH and pOH of a solution whose (H^+) ion concentration is 3.0×10^{-2} moles/ litre?

Solution:

$$pH = -\log [H^+]$$
$$= -\log (3 \times 10^{-2})$$
$$= -(\log 3 - \log 10^{-2})$$
$$= -[0.477 - 2]$$
$$= [-0.477 + 2]$$
$$pH = 1.523$$

As $pH + pOH = 14$

$1.523 + pOH = 14$

$pOH = 14 - 1.523$

$pOH = 12.477$

$pOH = 14 - 1.523$

$pOH = 12.477$

pH = 1.523 pOH = 12.477 (Answer)

Example. (4) Calculate the (H^+) ion concentration of a solution whose pH 4.4?

Solution:

$$pH = -\log [H^+]$$
$$pH = -\log (H^+) = 4.4$$
$$pH = \log (H^+) = -4.4$$

As we know that exponent should be whole number, so that -4.4 is equal to $(-5+0.6)$, then we can write as:

$$\log [H^+] = -4.4$$
$$\log [H^+] = (0.6-5)$$
$$H^+ = \text{Antilog } (0.6 - 5)$$
$$H^+ = \text{Antilog of } 0.6 \times \text{Antilog of } -5$$
$$= 3.98 \times 10^{-5}$$

$H^+ = 3.98 \times 10^{-5}$ M (Answer)

9.5.3 The Measurement of pH with (pH Paper):

There are three methods to measure the pH of a solution. (i) By acid-base indicator's . (ii) By "pH-meters" and (iii) By "pH-paper". The method for the approximate determination of pH widely used is with the 'pH-paper'. In this method, paper-strips, that are treated with several different indicators can be used to estimate pH. These papers-strips are called **"pH-paper"**. pH can be estimated by dipping the pH paper in a given solution, then by matching the colour appearing on the pH-paper with a colour corresponding to a known pH.

9.5.4 The Importance of pH:

The concept of pH plays essential role in the field of biology. Other areas in which pH information and control is necessary, include, water treatment, soil conditioning, swimming pool managements, corrosion, control, food processing and electroplating.

For example: The pH of human blood is normally maintained by the body between 7.35 and 7.45. If the blood pH drops to 7, as in some illness, the patient may go into coma, a pH below 6, death may occur. pH rises as high as 7.7 or 7.8 causes diabetes excess vomiting, diarrhea.

Hence the pH values of various body fluids are vey important for a doctor in diagnosing and treating many illnesses.

The pH values of several biological fluids.

Fluid	pH
Lemon juice	2.3
Vinegar	2.8
Tomato juice	4.2
Human urine	5.0-7.0
Cow's milk	6.5
Saliva	7.0
Human blood	7.35-7.45
Egg white	7.8

9.6 ACID-BASE TITRATION

There are two ways to make a solution of known molarity. The first and most convenient way to make the solution, by dissolving exactly 1

mole of solute in a litre (dm^3) of solution. The second way is to make up a solution quickly, using an estimated amount of solute and an estimated amount of solution, and then determine the solution's exact molarity by titration.

Titration is the chemical process by which we can determine the concentration of unknown solution, that reacts with a standard solution, whose concentration is known.

Titrations play a major role in determining, amounts of solutes present in a solutions. Titration process is an important tool of the analytical chemist.

9.6.1 Molarity (M) and Molar Solution:

The most generally useful means of expressing a concentration of solution is molarity.

It is defined as the number of moles of solute dissolved per 1 litre ($1dm^3$) of a solution, it is denoted by (M).

Thus 1 mole of H_2SO_4 (i.e. its gram formula mass) 98g dissolved in one litre ($1dm^3$) of solution is said to be 1 molar (1M) solution, if only one-half the mole i.e. 49g of H_2SO_4 are dissolved in one litre of solution, the solution is said to be one-half molar (i.e. 0.5 M)

The molarity of any solution is found by dividing the number of moles of solute by the number of moles/litres of the solution.

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Litres of solution}}$$

We known that,

(i) $$\text{Number of moles} = \frac{\text{Mass of solute}}{\text{Gram formula mass}}$$

(ii) $$\text{Litres of solution} = \frac{\text{Volume of solution in cm}^3\text{(ml)}}{1000}$$

$$\text{Molarity} = \text{(M)} = \frac{\text{Mass of solute}}{\text{Gram formula mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3\text{(ml)}}$$

For example:

(i) Calculate the molarity of a solution, containing 1.5g of NaOH in 250 cm^3 of solution?

Solution:

$$M = \frac{\text{Mass of solute}}{\text{Gram formula mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

Data:

(i)	Molarity	=	M = ?
(ii)	Mass of solute (NaOH)	=	1.5g
(iii)	Gram formula mass of solute (23+16+1)	=	40g
(iv)	Volume of solution	=	250 cm^3

$$M = \frac{1.5g \times 1000}{40 \times 250} = 0.15M$$

Result: Molarity of NaOH solution is = M = 0.15M

Example. 2

(i) What mass of(NaOH) must be dissolved in 500cm^3 of solution to make 1.5 M-solution?

Solution:

$$M = \frac{\text{Mass of solute}}{\text{Gram formula mass of solute}} \times \frac{1000}{\text{cm}^3 \text{ of solution}}$$

By cross-multiplication, we get,

$$\text{Mass of solute} = \frac{\text{M x G.F.mass x cm}^3 \text{of solution}}{1000}$$

Data:

(i)	Mass of solute (NaOH)		=	?
(ii)	Molarity	= M	=	1.5 M
(iii)	G.F mass of solute (23+16+1)		=	40 g
(vi)	Cm^3 of solution		=	500 cm^3

$$\text{Mass of solute} = \frac{1.5g \times 40 \times 500cm^3}{1000\ cm^3} = 30g$$

Result: Mass of solute i.e. NaOH = 30g

9.6.2 Preparation of Solution of Known Molarity:

For example, how to prepare 1 M-solution of NaOH.

To prepare 1M solution of NaOH you first weigh out one mole of NaOH i.e. 23 + 16 + 1 = 40g of NaOH, place 40g of NaOH in a 1 litre (1 dm^3) volumetric flask, and add some water to bring the level of solution to

calibration mark as shown in figure given below, finally the solution is shaken untill uniform.

Marker showing known volume of solution

Meniscus

(a) (b) (c)

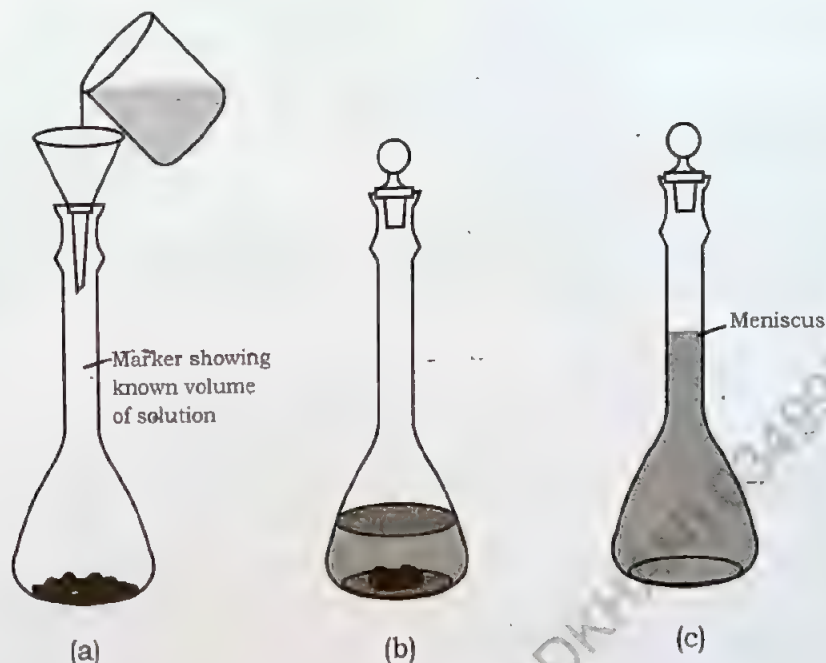



Fig. 9.2 Preparation of solution of known molarity

9.6.3 Standard Solution:

A solution whose molarity or strength is known, is called standard solution. For example 1M.KOH solution, contains 1mole of (KOH) i.e. formula mass expressed in grams =39+16+1 = 56g of KOH are dissolved in 1 litre of solution is said to be 1 molar (1M). If we use only one-half mole i.e. 28g of KOH in 1-litre of solution our solution would be one-half molar i.e. 0.5 M.

9.6.4 Acid-Base Titration:

In acid-base titration, a solution of known concentration (say base) is added gradually to a solution of unknown concentration (say an acid) so as to determine the concentration of unknown solution. The point at which the reaction is completed is called the end point.

Steps for Carrying out Titration:

1. Fill the burette with the given solution of NaOH; read and record the initial reading of the burette (Remember to read the lower meniscus) if the burette is not full refill it with NaOH solution. Then let the burette drain into beaker down to the zero mark before using it. The solution in the burette is called **"Titre".**

2. Pipette out 10 cm^3 of HCl in a conical flask, (titration flask) and add one or two drops phenolphthalein indicator. The solution in titration flask is called **"Titrant"**. A suitable acid-base indicator, such a methyl orange, litmus, or phenolphthalein is used to detect the end point.

S.No.	Indicator	Colour in acids	pH-range	Colour in bases
1.	Methyl orange	Red	3 - 5	Yellow
2.	Litmus	Red	6 - 8	Blue
3.	Phenolphthalein	Colourless	8 -10	Pink(red)

3. Add slowly the NaOH solution, from the burette into the conical flask (titration flask) with constant shaking. Stop adding the NaOH solution, when the mixture in the titration flask becomes light pink. Record the final burette reading. If you have been adding the reactant rapidly and think you have over run the end point, repeat the entire titration slowing down when you think you are near the end point. Rinse out the flask with water before the next sample is titrated.

4. Record the initial and final burette readings for each experiment.

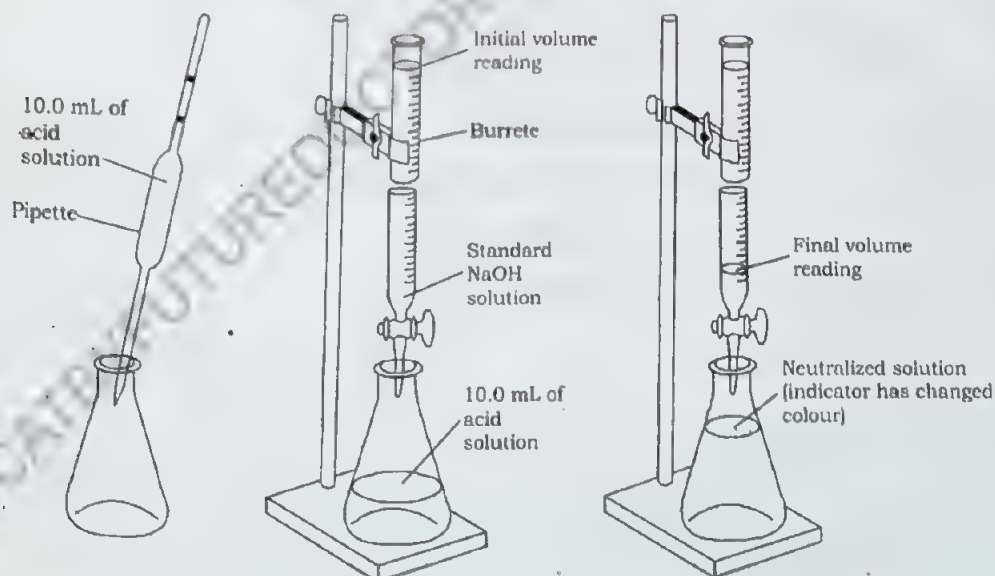



Fig. 9.3 Shows procedure for titrating an acid against a standardized solution of NaOH (base)

To see how titration works, let us consider that we have (HCl) solution (an acid) whose concentration we want to find by allowing it to react with a standard base such as NaOH.

We begin the titration by measuring out a known volume of HCl in a conical flask (titration flask)and adding few drops of an indicator, such as phenolphthalein (colourless liquid in acidic solutions but turns pink colour in basic solutions).

Next, we fill a burette with the(NaOH) standard solution (of known concentration) and we slowly add the NaOH to the HCl, until the phenolphthalein just begins to turn pink, indicating that all the HCl has been reacted and that the solution is starting to become basic. Then by recording the reading from the burette, to find the volume of the NaOH (standard solution) that has reacted with a known volume of HCl, we can calculate the concentration (molarity) of the HCl.

Let us assume for example that we take 10 mls of HCl-solution and find that we have to add 30 mls of 0.1 M NaOH from the burette to obtain complete reaction. Using the following formula we can calculate the molarity of unknown solution of HCl?

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

The balanced equation shows that:

1 mole of HCl reacts with 1 mole of NaOH

1-HCl ⟶ 1-NaOH

Burette = NaOH solution molarity = M = 0.1 M.

Pipette = HCl solution molarity = M = ?

Indicator = Phenolphthalein.

Colour change = Colourless to pink.

V_1 = Volume of HCl taken = 10 mls

M_1 = Molarity of HCl taken = ?

V_2 = Volume of NaOH solution = 30 mls

M_2 = Molarity of NaOH solution = 0.1M

n_1 = Number of moles of HCl = 1

n_2 = Number of moles of NaOH = 1

$$\frac{V_1 \times M_1}{n_1} = \frac{V_2 \times M_2}{n_2}$$

$$\frac{10 \times ?}{1} = \frac{30 \times 0.1}{1} = \frac{30 \times 0.1}{10} = \frac{3}{10}$$

$M_1 = .3M$

Molarity of HCl = 0.3 M

We can also find out the amount of HCl dissolved per 1-litre = ($1dm^3$) of a solution by the following formula.

Amount of HCl per dm^3 = molarity x F. mass in grams
Amount of HCl per dm^3 = 0.3 x 36.5 = 10.95g
(As F. mass of HCl = 1+35.5 = 36.5g)

Result:	Molarity of HCl = 0.3M Amount of HCl/dm^3 = 10.95g

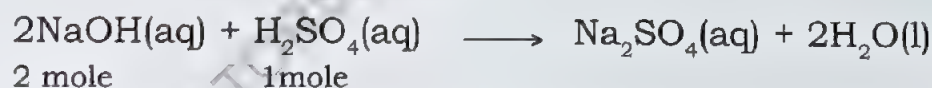
Example.

A flask contain 30 ml of NaOH solution, it require 50ml of 0.15M-H_2SO_4, to complete the reaction. Calculate the molarity of NaOH and how many grams of NaOH were in the flask?

Solution.

The balance equation for the reaction is:

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

2 mole 1mole

Since 2 mole of NaOH $\equiv$ 1mole of H_2SO_4 we need twice the amount of NaOH to react completely with H_2SO_4 solution as needed for the same concentration of HCl solution.

$$\frac{M_1 \times V_1}{n_1} = \frac{M_2 \times V_2}{n_2}$$

Where:

V_1 = Volume of H_2SO_4 = 50 ml
M_1 = Molarity of H_2SO_4 = 0.15M
n_1 = Number of moles of H_2SO_4 = 1
V_2 = Volume of NaOH = 30 ml
M_2 = Molarity of NaOH = ?
n_2 = Number of moles of NaOH = 2

$$\therefore \quad \frac{M_1 \times V_1}{n_1} = \frac{M_2 \times V_2}{n_2} \qquad \frac{0.15 \times 50}{1} = \frac{M_2 \times 30}{2}$$

$$M_2 = \frac{0.15 \times 50 \times 2}{30} = 0.5\ M$$

$$\text{Amount of NaOH per 30 ml} = \frac{M \times \text{G.F.mass} \times \text{volume of solution}}{1000}$$

$$= \frac{0.5 \times 40 \times 30}{1000} = 0.6g$$

Result:	Molarity of NaOH = 0.5M Amount of NaOH in the flask = 0.6g

SUMMARY

(1) Acids are substances that have the following properties in aqueous solution.

- (i) They have a sour taste.
- (ii) Change the colour of litmus from blue to red.
- (iii) React with active metals (such as iron, tin and zinc) to liberate hydrogen gas.
- (iv) React with bases to form salt (ionic compounds) and water.
- (v) Their aqueous solutions conduct an electric current because they contain ions, they are electrolytes.

(2) Similarly bases are defined as substances that have the following properties, when dissolved in water.

- (i) They have bitter taste.
- (ii) Feel soapy and slippery on the skin.
- (iii) Turn colour of litmus paper from red to blue.
- (iv) React with acids to form salt and water.
- (v) They give aqueous solutions which conduct an electricity, they are electrolytes.

(3) Arrhenius proposed that an acid is a substance that yields hydrogen (H^+) ions in water solution, while base produces hydroxide (OH^-) ions.

(4) The Lewis theory defines a base, as any substance that has one or more pair of electrons for bond formation. A Lewis base is an electron-pair donor, while lewis acid is electron-pair acceptor.

(5) Strong acid is one that ionizes in dilute solution almost completely and weak acid is one that ionizes only slightly. Thus a solution of strong acid has relatively high concentration of (H^+) ions and a solution of weak acid has a relatively low concentration of (H^+)ions. Similarly strong base is one that ionizes in dilute solution almost completely and weak base is one which ionizes only slightly. Thus a solution of strong base has a relatively high concentration of (OH^-) ions and a solution of weak base has a relatively low concentration of (OH^-) ions.

(6) Acids may be classified in terms of number of proton per molecule of acid that can given up in a reaction. Acids $HClO_4$, HI, HBr, HCl, HNO_3 and HCN that contain one ionizable hydrogen atom in a molecule of acid are called monoprotic acids. Di-protic acids contain two ionizable hydrogen atoms in one molecule of the acid, ionization of such acids occurs in two steps. Tri-protic acids, such as phosphoric acid, ionizes in three steps. Mono protic, Di-protic and tri-protic acids are commonly called mono basic, dibasic and tribasic. Similarly, bases that produces one (OH^-) ion per molecule are called mono-acidic bases, that produces two (OH^-) ions per molecule are called di-acidic bases and that produce three (OH^-) ions per molecule are called tri-acidic bases.

(7) When acids react with bases, salt and water are formed, and is known as neutralization. Many of these reactions can be denoted by a single net-ionic equation.

$$H^+ + OH^- \longrightarrow H_2O.$$

Salts are formed by any positive ion except hydrogen, combined with any negative ion, except hydroxide ion. Salts are ionic substances which are completely dissociated in water solution and are known as electrolytes.

(8) The acidity of an aqueous solution is expressed by its pH, which is defined as the negative logarthim of the hydrogen ions concentration. (in moles per litre). The neutral solution has pH=7, acidic solution has a pH is less then 7 and basic solution has pH is more then 7. Nowadays pH of a solution is determined by pH-meter.

(9) Most salts are strong electrolytes and dissociate completely into ions in aqueous solutions. The reaction of these ions, with water lead to the acidic or basic solution, called hydrolysis, the solution of strong acids and weak bases are acidic and the solutions of weak acids and strong bases are basic in nature, but the solutions of strong acids and strong bases are neutral because of the formation of equal concentration of strong acid and strong base.

(10) Water is amphoteric because it is both an acid and a base. It undergoes self ionization to give small quantities of hydrogen ion (H^+) and hydroxide ion (OH^-).

(11) Standard solution is that whose concentration is known.

(12) The concentration of solution is usually expressed as molarity (M), which is defined as the number of moles of solute dissolved per litre of solution. A solution of exact concentration can often be determined by titration. Titration is a chemical process by which we can determine the concentration of unknown (acid) solution by reacting it with a solution of known concentration of (base) solution.

EXERCISE

1. Fill in the blank.

i. The formula of baking soda is ....

ii. The formula of epsom salt is ....

iii. $K_2SO_4Al_2(SO_4)_3 . 24H_2O$ is the formula of ....

iv. .... is the most convenient way of expressing concentration.

v. The molarity of solution is denoted by ....

vi. A solution whose strength is known is called ....

vii. If H^+ ion concentration of a solution is 1×10^{-14} M, the solution is ....

viii. If the OH^- ion concentration of a solution is 1×10^{-10}M, solution is............

ix. .... is the process by which we can determine the concentration of un-known solution with the help of standard solution.

x. The solution whose H^+ ion concentration is 1×10^{-4} M, then its pH is ....

xi. The solution whose pH is 6, then its H^+ ion concentration is ....

xii. The volume of a pipette is generally .... ml or cm^3.

xiii. Molarity is defined as number of moles per ....

2. Tick the Correct Answer:

(i) The substances whose aqueous solution change the blue litmus to red:

(a) Acids. (b) Bases.
(c) Neutral. (d) Salts.

(ii) The substances having a tendency to lose one or more protons are called:

(a) Acids. (b) Bases.
(c) Neutral. (d) Salts.

(iii) The substances which donate the pair of electrons for bond formation, are known as:

(a) Acids. (b) Bases.
(c) Neutral. (d) Salts.

(iv) When equivalent quantities of acid and base are mixed, salt and water are formed, the reaction is termed as:

(a) Hydration. (b) Hydrolysis.
(c) Neutralization. (d) None of these.

(v) The acids which contain one acidic hydrogen are called:

(a) Mono-protic. (b) Di-protic.
(c) Tri-protic. (d) Poly protic.

(vi) The number of acidic hydrogen atoms present in a molecule of an acid is called:

(a) Acidity. (b) Basicity.
(c) Neutral. (d) Hydrolysis.

(vii) The number of replaceable $[OH]^-$ ions present in a molecule of base, is called:

(a) Acidity. (b) Basicity.
(c) Neutral. (d) Hydrolysis.

(viii) An acid that produces large number of (H^+) ions in aqueous solution is called:

(a) Strong base. (b) Weak base.
(c) Strong acid. (d) Weak acid.

(ix) An ionic compound, that is formed when an acid neutralizes a base, is called:

(a) Acids. (b) Bases.
(c) Neutral. (d) Salts.

(x) Salts that formed by the reaction of strong acid with weak base are:

(a) Neutral. (b) Acidic.
(c) Basic. (d) Normal.

(xi) Salts that formed by the reaction of weak acid with strong base are:

(a) Acidic. b) Basic.
(c) Neutral. d) Normal.

(xii) Alums are:

(a) Single salts. b) Double salts
(c) Triple salts. d) Normal salts

xiii. The formula of Washing Soda is:

(a) Na_2CO_3. (b) $Na_2CO_3 . 6H_2O$.
(c) $Na_2CO_3 . 10H_2O$. (d) $NaHCO_3$.

3. Write answer of the following questions:

(i) What is Arrhenius theory of acids and bases? Why is the Arrhenius theory not satisfactory for acids and bases?

(ii) What is Lewis theory of acids and bases?

(iii) List the main general properties of acids and bases.

(iv) Write the formulas of four strong acids and four weak acids.

(v) Sulphuric acid (H_2SO_4) is strong acid, (HSO_4^-) is a weak acid? Account for the difference in strength?

(vi) Explain with illustrations what are strong acids and bases and weak acid and bases?

(vii) (a) Define acidic, basic and neutral solutions in terms of (H^+) ion concentration. Indicate whether each of the following solution will be acidic, basic or neutral.

(b) Strong acid and strong base.
(c) Strong acid and weak base.
(d) Weak acid and strong base.

(viii) What is salt? Give four examples of Salt?

(ix) Give an example each of mono-protic acid, di-protic acid and a tri-protic acid?

(x) Identify the following as a weak or strong acids or bases?

(a)	NH_3	(b)	H_3PO_4	(c)	LiOH
(d)	HCOOH (formic acid)			(e)	H_2SO_4
(f)	H_2CO_3	(g)	$Ba(OH)_2$		

(xi) Define molarity? What is molar solution?

(xii) What is molarity of H_2SO_4 solution containing 9.8g of H_2SO_4 per 500 ml?

(xiii) Define pH? Explain why a neutral solution is said to have a pH of seven.

(xiv) Give an equation to show the dissociation of water.

(xv) Define the term "Amphoteric". Give an example?

(xvi) Calculate the pH of the following solutions.
(a) .001 M-HCl (b) 5.2×10^{-4} M-HNO_3

(xvii) Calculate (H^+) ion concentration of solutions, whose

(a) pH = 5.2 (b) pH = 9.63

(xviii) Fill in the word "acidic" "basic" or "neutral" for the following solutions.

(a) pH = 7 ....

(b) pH = > 7 ....

(c) pH = < 7 ....

(xix) The pOH of a solution is 9.40? Calculate the (H^+) ion concentration.

(xx) Define acid-base neutralization. Does the definition imply that the resulting solution is always neutral?

(xxi) Describe clearly how a solution of HCl could be titrated with a solution of NaOH?

(xxii) Define acid-base titrations, standard solution and equivalence point?

(xxiii) What volume of 0.5M KOH solution is needed to neutralize completely in each of the following.

(a) 10.0 ml of 0.3 M HCl-Solution.

(b) 10.0 ml of 0.2 M H_2SO_4-Solution.

(c) 10.0 ml of 0.25M H_3PO_4-Solution.

CHAPTER-10

CHEMICAL ENERGETICS

You will learn in this chapter about :

* Exothermic and Endothermic reaction.
* Heat of Reactions.
* Measurement of heat of reactions.
* Heat of Neutralization.

INTRODUCTION

The chemical reactions during which material changes are accompained with change in heat energy are called thermochemical reactions. The branch of chemistry which deals with the study of heat changes in chemical reactions is called thermochemistry. There are two types of thermochemical reactions i.e. exothermic and endothermic chemical reactions.

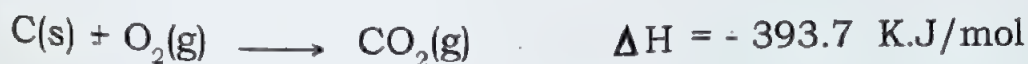
10.1 EXOTHERMIC AND ENDOTHERMIC REACTIONS

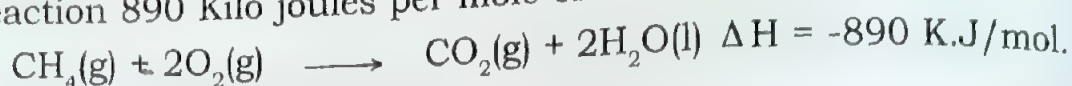
10.1.1 Exothermic Reaction:

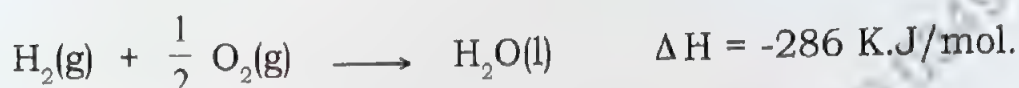
It is Greek word and composed of EXO means out of or to evolve + THERME means heat. An Exothermic Reaction is the chemical change during which heat is given out or released. The change of heat is represented by ΔH and it is shown by negative sign.

Examples:

(i) The combustion of coal in air is the example of exothermic reaction. 393.7 Kilo joules of heat energy is released when 1 mole of coal is burnt in 1 mole of O_2 to produce 1 mole of CO_2.

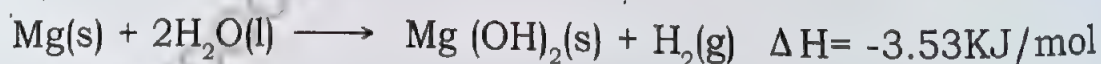
$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.7 \text{ K.J/mol}$$

(ii) Burning of methane in presence of oxygen is another example of exothermic reaction. When 1 mole of methane is burnt in 2 moles of O_2 then 1mole of CO_2 and 2 moles of water are formed. During this reaction 890 Kilo joules per mole of heat energy is released.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890\ K.J/mol.$$

(iii) The formation of water from hydrogen and oxygen is also example of exothermic reaction. 286 Kilo joules per mole of heat energy is released, when 1mole of H_2 reacts with $\frac{1}{2}$ moles of O_2 to form 1 mole of H_2O.

$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l) \qquad \Delta H = -286\ K.J/mol.$$

Generally in exothermic reactions, heat flows from the system to surroundings and container becomes hot, this is only possible, when total energy of the reactants is greater than total energy of products. The difference in the energies is the heat supplied by the system to surroundings.

Using Exothermic Reactions to Warm Food:

In modern army, food rations can be warmed without benefit of stove or camfire. The pouch that contains the food is attached to flameless radiation heater. The heater contains chemicals that react with water to produce heat. When the pouch is placed in a bag and water added, temperature of the food reaches to 60^0C in about 15 minutes.

$$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(s) + H_2(g) \quad \Delta H = -3.53KJ/mol$$

The reaction of (Mg) with water is slow, because of the formation of film oxide (MgO). The reaction of Mg with water is highly accelerated in the presence of iron (Fe) and ordinary salt (NaCl). Thus, the flameless radiation heater contains a mixture of Mg, Fe and NaCl.

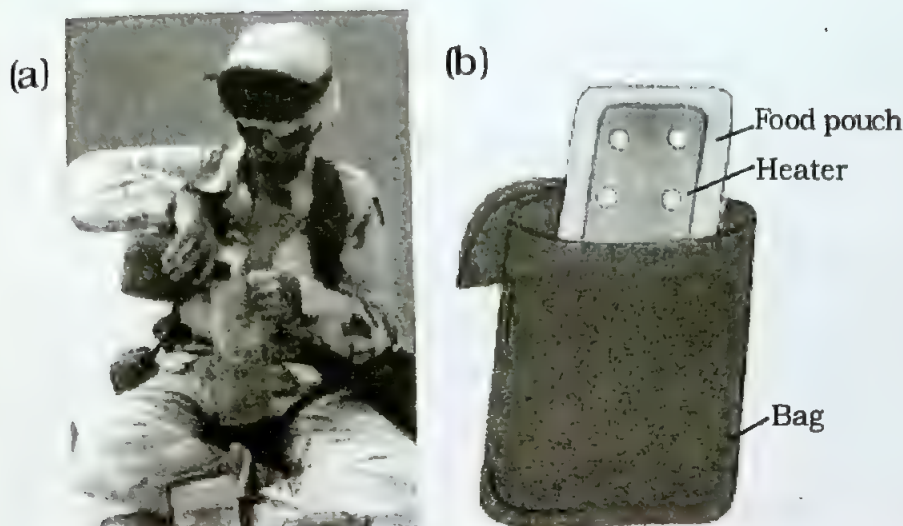



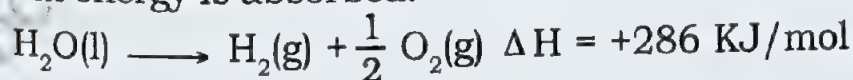
Fig. 10.1(a) A soldier adding water to the bag containing pouch
Fig. 10.1(b) The food pouch is attached with chemical heater containing Mg, Fe and NaCl.

10.1.2 Endothermic Reactions:

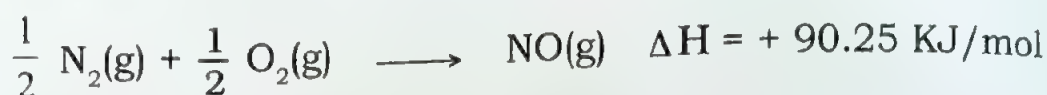
It is Greek word and composed of ENDO mean into or to absorb and THERME means heat. Endothermic reaction is the chemical change during which heat is absorbed or taken in. The change of heat energy is represented by ΔH and sign of ΔH is positive ($+\Delta H$= Absorbs heat). During the endothermic reaction heat is absorbed from the surroundings, it means heat flows from surroundings to the system and container becomes cold. This is because the total energy of the products is greater than the total energy of reactants. Here, the difference in the energies is the heat supplied to the system by the surroundings.

Examples:

(i) The decomposition of water into hydrogen and oxygen is example of endothermic reaction. During decomposition of 1mole of water to 1 mole of hydrogen and half mole of oxygen 286 Kilo joules per mole of heat energy is absorbed.

$$H_2O(l) \longrightarrow H_2(g) + \frac{1}{2} O_2(g) \quad \Delta H = +286 \text{ KJ/mol}$$

(ii) 1 mole of nitric oxide (NO) is formed by combination of $\frac{1}{2}$ mole of N_2

and $\frac{1}{2}$ mole of O_2. This is the example of endothermic reaction and heat absorbed is about 90.25 Kilo joules per mole.

$$\frac{1}{2} N_2(g) + \frac{1}{2} O_2(g) \longrightarrow NO(g) \quad \Delta H = + 90.25 \text{ KJ/mol}$$

10.2 HEAT CONTENTS OF REACTION

Every substance possesses a characteristic internal energy and internal energy depends upon the structure and physical state of that substance. The energy possessed by a substance is called heat contents of that substance. During a chemical reaction, the reactants are converted into products and heat energy is either absorbed or evolved. This is because the heat contents of these respective substances are different. The heat evolved or absorbed at constant pressure is called as enthalpy of the reaction.

The heat content (enthalpy) of a substance is represented by "H" and the change in heat content during a chemical reaction is then represented by ΔH. Here Greek letter Δ (delta) signifying the change in the property. It means the change in heat content during chemical reaction is the difference between the heat content of products and reactants of that reaction.

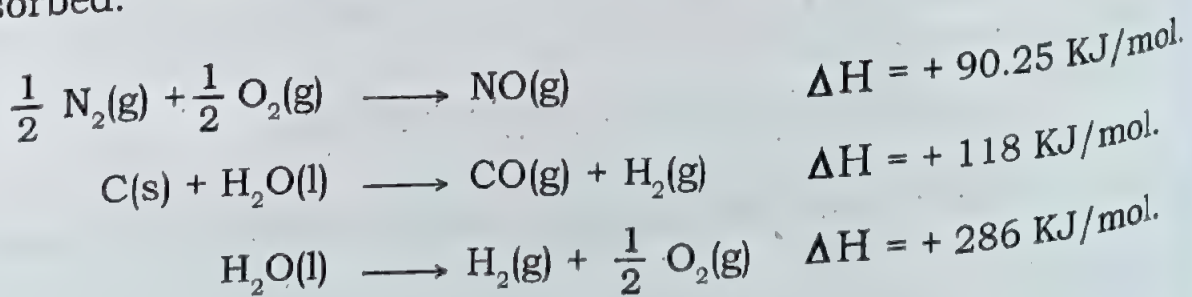
It is difficult to measure the enthalpy of a reaction, but we can measure the change in enthalpy which is denoted by (ΔH). It is obtained by subtracting the enthalpy of reactant (H_1) from enthalpy of products (H_2).

$$\underset{\text{Change of Enthalpy}}{\Delta H} = (\underset{\text{Heat content of product}}{H_2} - \underset{\text{Heat content of reactants}}{H_1})$$

If enthalpy of products is greater than the enthalpy of reactants, then the sign of ΔH will be positive and over all reaction is endothermic and heat is absorbed.

$$\frac{1}{2} N_2(g) + \frac{1}{2} O_2(g) \longrightarrow NO(g) \quad \Delta H = + 90.25 \text{ KJ/mol.}$$

$$C(s) + H_2O(l) \longrightarrow CO(g) + H_2(g) \quad \Delta H = + 118 \text{ KJ/mol.}$$

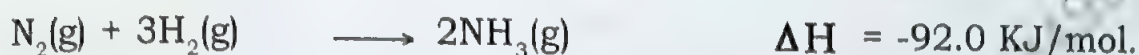
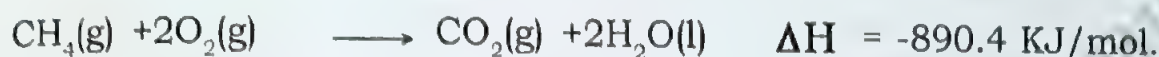
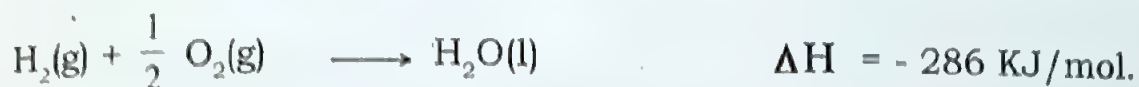
$$H_2O(l) \longrightarrow H_2(g) + \frac{1}{2} O_2(g) \quad \Delta H = + 286 \text{ KJ/mol.}$$

If enthalpy of product is smaller than the reactants, then the sign of ΔH will be negative and overall reaction is exothermic and heat is evolved.

The examples of exothermic reactions are following:

$$H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(l) \qquad \Delta H = -286 \text{ KJ/mol.}$$

$$S(g) + O_2(g) \longrightarrow SO_2(g) \qquad \Delta H = -296.8 \text{ KJ/mol.}$$

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = -890.4 \text{ KJ/mol.}$$

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g) \qquad \Delta H = -92.0 \text{ KJ/mol.}$$

10.3 MEASUREMENT OF HEAT OF REACTION

The heat absorbed or evolved during thermochemical reaction is called heat of reaction. Exothermic and endothermic reactions can eaisly be detected by touching the vessel before and after chemical reaction. The increase in temperature indicate that reaction is exothermic and decrease in temperature indicates that the reaction is endothermic. The accurate values of ΔH can be determined by using calorimeter. The simple type of calorimeter is an insulated container fitted with thermometer and a stirrer. Known amounts of reactants are placed in calorimeter, when reaction proceeds the heat energy evolved or absorbed will either cool or warm the system. ΔH for reaction may be calculated by determining the difference in temperature, mass of reactants and specific heat of reaction mixture.

Heat of Neutralization:

The reaction between an acid and base to form a salt and water is called neutralization reaction. Neutralization reaction is an example of exothermic reaction. The amount of heat released during a neutralization reaction in which 1 mole of water is formed is called as the heat of neutralization or the amount of heat released when 1 mole of hydrogen ions (H^+) from an acid reacts with 1 mole of hydroxide ions (OH^-) from base to form salt and one mole of water is called heat of neutralization.

Procedure:

Take 50 cm^3 of molar NaOH solution and note its temperature (t_1) and 50 cm^3 of molar HCl solution and note its temperature (t_2). The two temperatures will be usually same, but they need not. Pour the HCl solution in 250 mls beaker (Calorimeter) and then add quickly NaOH solution the solution being stirred all the time, and note down the highest temperature

reached during the reaction. At the end weigh the calorimeter with salt solution. Heat of neutralization is calculated by the following formula.

$$\Delta H = m \times S \times (t_2 - t_1)$$

or $$\Delta H = m \times S \times \Delta t$$

Observation and Calculations

1. Mass of calorimeter along with stirrer $= w_1 = 50g$
2. Mass of calorimeter along with stirrer + salt sol. $= w_2 = 150g$
3. Mass of salt solution $(w_2 - w_1)$ $= m = 100g$
4. Specific heat of salt solution $= s = 4.25 J/g/^0C$
5. Initial temperature of reactants $t_1 = t_2$ $= t^0C = 20^0C$
6. Final highest temperature $= t_3{}^0C = 26.8^0C$
7. Increase in temperature i.e. $t_3{}^0C - t^0C$ $= \Delta t = 6.8^0C$

Then

Heat of neutralization is given by $\Delta H = ms\,\Delta t$

$$\Delta H = 100g \times \frac{4.25J}{1g \times 1^0C} \times 6.8^0C = 2856J$$

The value obtained i.e 2856J is for 50 mls of solution, it must be multiplied by 20 to give the amount of heat evolved, when 1 mole of NaOH is neutralized by HCl

$\therefore$ $2856\,J \times 20 = 57120\,J = 57.12\,K.J/mol$

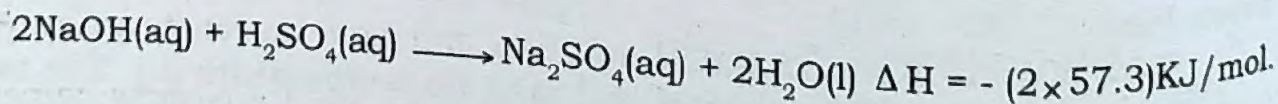
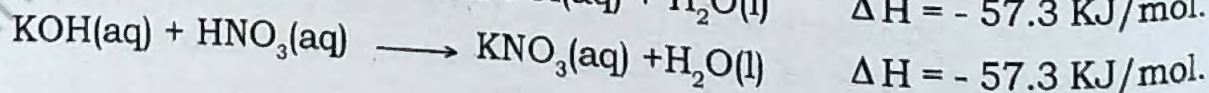
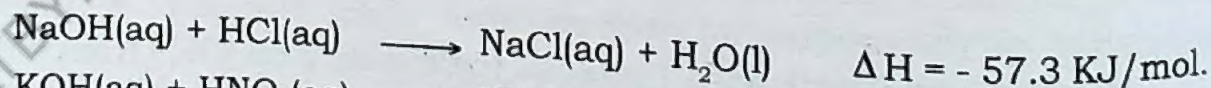
Result: The heat of neutralization of NaOH by HCl is $\Delta H = -57.12\,KJ/mol$.

Note: The heat of neutralizaiton for any strong acid with strong base is approximately same.

For example:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) \quad \Delta H = -57.3\,KJ/mol.$$

$$KOH(aq) + HNO_3(aq) \longrightarrow KNO_3(aq) + H_2O(l) \quad \Delta H = -57.3\,KJ/mol.$$

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l) \quad \Delta H = -(2 \times 57.3)KJ/mol.$$

SUMMARY

1. Heat energy is given out in exothermic reaction.
2. Heat energy is absorbed in endothermic reaction.
3. In chemical reactions, the material change is accompained with the change of heat energy.
4. There are two types of thermochemical reactions i.e. exothermic and endothermic reaction.
5. All combustion reactions are the examples of exothermic reactions.
6. The enthalpy is represented by "H" and change of enthalpy is represented by "ΔH".
7. The energy given out or absorbed at constant pressure is called enthalpy of the reaction.
8. Another name for heat content of a reaction is enthalpy of the reaction.
9. In neutralization reactions, the heat is evolved, which is called as heat of neutralization.
10. The change in heat contents during a chemical reaction is the difference between the heat content of products and reactants.
11. When enthalpy of products is greater than the enthalpy of reactants than the reaction will be endothermic.
12. The sign of ΔH for exothermic reactions is negative (-ve)
13. The sign ΔH for endothermic reaction is positive. (+ve)

EXERCISE

1. **Fill in the blanks:**
   (i) The heat given out in a chemical reaction is called .... reaction.
   (ii) In .... reaction, heat is taken in.
   (iii) Heat evolved or absorbed during chemical reaction at constant pressure is called ....
   (iv) Acid base reaction is called .... reaction.
   (v) $C(s) + O_2(g) \longrightarrow CO_2(g)\ \Delta H$ = ------ KJ/mol.
   (vi) $H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(l)\ \Delta H$ = ------ KJ/mol.

2. **Tick the correct answers:**
   (i) In an exothermic reaction.
      (a) Heat energy is lost. (b) Heat energy is gained.
      (c) Heat energy is lost as well as gained.
      (d) None of them.

(ii) In an exothermic reaction.
  (a) Container becomes hot.
  (b) Container becomes cold.
  (c) The temperature of container remains the same.
  (d) None of them.

(iii) During an endothermic reaction.
  (a) Container used becomes cold.
  (b) Container used becomes hot.
  (c) The temperature of container used remains same.
  (d) Total energy of reactants increases.

(iv) The heat evolved during the formation of 1 mole of water from H_2 and O_2 is
  (a) 286 Kilo joules/mol (b) 186 Kilo joules/mol
  (c) 300 Kilo joules/mol (d) 200 Kilo joules/mol

(v) The formation of water form H_2 and O_2 is example of,
  (a) Exothermic reaction. (b) Endothermic reaction.
  (c) Neutralization reaction. (d) None of them.

3. Write answer of the following questions:

(i) Define the following terms:
  (a) Thermo chemistry. (b) Exothermic reaction.
  (c) Endothermic reaction

(ii) Give atleast two examples of exothermic reactions and two examples of endothermic reactions.

(iii) Which of the following are exothermic or endothermic processes?
  (a) The decomposition of mercuric oxide (HgO).
  (b) The electrolysis of water.
  (c) The reaction of (Na) with water.
  (d) The burning of methane CH_4.
  (e) The decomposition of $KClO_3$.
  (f) A match burn.

(iv) Define the following terms:
  (a) Enthalpy. (b) Enthalpy of reaction.

(v) Define heat of neutralization. What would be the value of heat of neutralization, when strong acid reacts with strong base?

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